







# SCIENCE OF COMMON LIFE





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# SCIENCE OF COMMON LIFE

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# PREFACE .

THE value of an introduction to the scientific method in the school training of both sexes is now fully recognised. There is still some difference of opinion, however, as to how girls can best be taught to appreciate the value of experiment and deduction in the endeavour to understand the forces of Nature. Fortunately, we believe, the idea gains ground that there are no more suitable means than the ordinary operations and practices of the household to illustrate physical and chemical laws. Adopting the general trend of the recent syllabus of the Board of Education, though in no sense slavishly following its details, we have attempted to make students familiar with those broad principles of science without which the intelligent guidance of the domestic economy is impossible. .

Adhering to the belief that it is only by experiment that the truths presented can be made real to young minds, we have arranged a large number of suitable practical exercises the due performance of which will provide a sure foundation for the later explanations of the teacher. Unless the experiments are carried out by, or at least in the presence of, the pupil, much of the value of the course of study will be lost. Since, too, our experience shows that as many of the faculties of the pupil as possible should be brought into play, special attention has been paid to the illustrations throughout the book. The majority of

them are new, and we are glad to express our indebtedness to Mr. O. L. Lacpur for the care he has taken in representing apparatus and experiments in an artistic and attractive manner.

We gratefully acknowledge how much the book owes to Prof. R. A. Gregory. Not only has he given us free permission to use any parts of previously published volumes written by him in association with one of us, but throughout the preparation of this volume he has placed his broad experience and judicious criticism at our disposal. We are sensible of the fact that any merits the following pages may prove to possess are in a large measure due to his generous assistance.

A. T. SIMMONS,  
E. STENHOUSE.

*September, 1901.*

## PREFACE TO SECOND EDITION

THE demand for a new edition of this book has provided the opportunity to change its title. Many teachers, more particularly perhaps those in schools for girls, are introducing simple experimental work designed to give their pupils sound general ideas of the scientific principles underlying everyday phenomena; and they are seeking a class-book arranged with this object in view. To direct the attention of teachers to such a book, while at the same time indicating that the volume provides an introduction to the principles of domestic science, the title *Science of Common Life* has been adopted instead of *Experimental Hygiene*, which has in some quarters given rise to ambiguity.

A. T. S.  
E. S.

*November, 1904.*

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# EXPERIMENTAL HYGIENE

## CHAPTER I

### LONG AND SQUARE MEASURE

#### MEASUREMENT OF LENGTH

1. **Inches and fractions.**—(a) *Eighths of an inch.* Obtain a tape measure. Notice that each inch is divided into a certain number of equal parts—called fractions of an inch. Usually there are eight of these parts marked on each inch of a tape measure; therefore there are four in half an inch and two in a quarter of an inch (Fig. 1). Write down

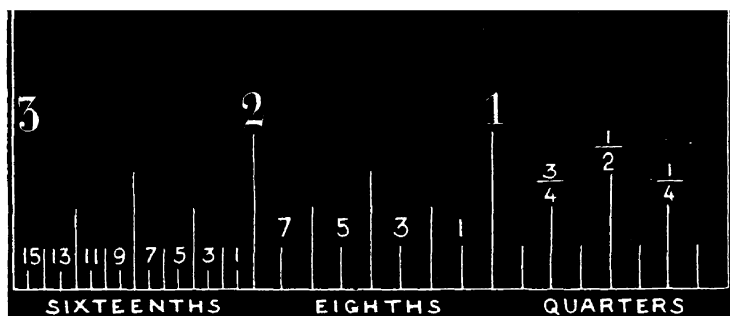


FIG. 1.—Fractions of an inch.

the following fractions of an inch in the order of their length, putting the greatest first and the smallest last:  $\frac{1}{2}$  in.,  $\frac{3}{8}$  in.,  $\frac{1}{4}$  in.,  $\frac{7}{8}$  in.,  $\frac{3}{4}$  in.,  $\frac{1}{8}$  in.,  $\frac{5}{8}$  in.

(b) *Sixteenths of an inch.* More accurate measures have each inch divided into sixteen instead of eight equal parts. If you have such a measure, find, by counting, the number of sixteenths of an inch there are in  $\frac{1}{4}$  in.,  $\frac{1}{2}$  in., and  $\frac{3}{4}$  in. If your inches are not divided into sixteenths, you should refer to Fig. 1, which shows inches divided



into this number of parts. Write down which is the greater fraction in each of the following pairs, referring to your measure, or Fig. 1, if necessary:  $\frac{1}{2}$  or  $\frac{9}{16}$ ,  $\frac{3}{8}$  or  $\frac{1}{4}$ ,  $\frac{2}{3}$  or  $\frac{1}{2}$ ,  $\frac{5}{8}$  or  $\frac{3}{4}$ ,  $\frac{1}{3}$  or  $\frac{1}{4}$ ,  $\frac{1}{2}$  or  $\frac{3}{8}$ ,  $\frac{5}{8}$  or  $\frac{3}{4}$ .

**2. Simple measurements.**—Using a tape measure, find the length, or width, of any convenient object, such as the top of a table or desk. Write down the length in feet, inches, and fractions of an inch, thus:

Width of desk, 2 feet,  $3\frac{3}{4}$  inches.

Length of sheet of foolscap paper, 1 foot,  $1\frac{1}{2}$  inches.

**3. Decimal fractions.**—Obtain a rule having the inches on one edge divided into ten equal parts instead of the eight or sixteen parts usually found (Fig. 2). How many tenths are there in half an inch? The tenths of an inch can be written like common fractions, for instance,  $\frac{1}{10}$  signifies one-tenth, and  $\frac{3}{10}$  means three-tenths. A more convenient way is to separate the inches from the tenths by means of a dot. Thus, a length of  $6\frac{3}{10}$  in. is written on this *Decimal System* as 6.3. The two ways of writing tenths in common and decimal fractions may be compared as follows:

Common }  $\frac{1}{10}$   $\frac{2}{10}$   $\frac{3}{10}$   $\frac{4}{10}$   $\frac{5}{10}$   $\frac{6}{10}$   $\frac{7}{10}$   $\frac{8}{10}$   $\frac{9}{10}$   $\frac{10}{10}$   
fractions }

Decimal } 0.1, 0.2, 0.3, 0.4, 0.5, 0.6, 0.7, 0.8, 0.9, 1  
fractions }

**4. Measurements in tenths.**—Measure one or two small objects as before, using your rule divided into inches and tenths instead of the tape measure. Write down the number of inches and tenths in the way described; thus, length of a new lead pencil, 7.1 inches.

**5. Metric measures of length.**—Notice that one edge of the rule provided is not divided into inches and fractions of inches, but is like the inside scale, B, shown in Fig. 2.

The smallest of these divisions are called *millimetres*. Ten millimetres make one of the numbered divisions, called *centimetres*; ten centimetres make one *decimetre*; and ten decimetres make one *metre*, which is a little longer than a yard.

**6. Relation between metric and British measures of length.**—Find by examining your rule the number of centimetres and tenths of a centimetre which are equal to the length of four inches. Find also the number of centimetres in  $7\frac{1}{2}$  or 7.5 inches,

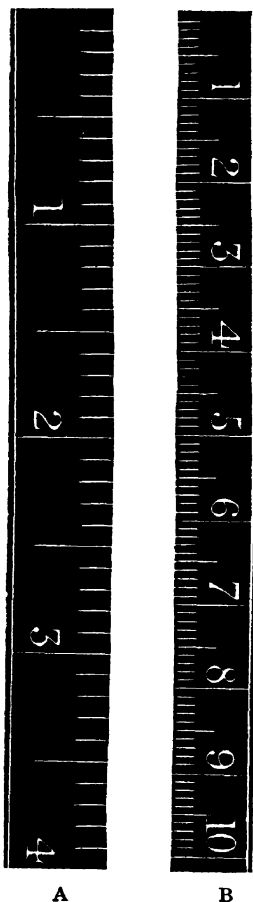


FIG. 2.—A, Inches and tenths. B, Centimetres and millimetres.

and in 11 inches, and the number of millimetres in the length of one inch. By comparing the two scales on your rule, observe that—

Number of centimetres in one foot = 30·5.

Number of inches in one decimetre = 3·9.

**7. Metre and yard.**—Upon the back of your tape measure, mark off a length equal to 100 centimetres, that is, 1 metre, starting from the point where the inches begin on the other side. Prick a hole through the measure at the point where you mark the length of a metre, then turn over the measure and notice where the hole occurs on the inch scale.

Number of inches in one metre = 39·3.

Number of centimetres in one yard = 91·4.

**Measurement of length.**—Every one is familiar with such common lengths as an inch, foot, and yard, and many people can make a fairly good estimate of any of these lengths, even if they have not a tape measure or rule with which to measure them. But, although it is possible to guess the length of a thing, a more satisfactory way is to measure it. In making such a measurement, a definite standard of length is used, represented, perhaps, by marks at a definite distance apart on a shop counter, or by lines marked upon a strip of wood or metal, or by a suitably divided ribbon of a material which is not liable to stretch. It is easy to understand that if a piece of string or ordinary tape of a yard in length were employed it could be stretched a little so as to be more than a yard long, and if it were wetted it would shrink to less than a yard long. It would, therefore, not be wise to use such a measure to determine the length of a thing. A yard, a foot, or an inch should be the same length always and everywhere, whether in a draper's shop or in a science laboratory. We know from common experience that, approximately, they are so. For the ordinary purposes of life, the tape measure and foot-rule are sufficiently accurate, but it is evident that in the course of time the manufacturers of such measures might make the lengths a little longer or shorter; or, at any rate, one of them might do so, and then declare his measures were the only true ones, while the other manufacturers held that his lengths were wrong. To secure uniformity, and provide a permanent length with which others may be compared if necessary, a standard of length has to be agreed upon.

**British standard of length.**—What is known as the

*Imperial standard yard* is the British standard unit of length. The yard is defined as "the length, at 62° F.,<sup>1</sup> marked on a bronze bar deposited with the Board of Trade."

The temperature of the bar has to be stated, because metals expand when they are heated and contract when cooled, so that the bar is slightly longer when it is warmer than sixty-two degrees Fahrenheit, and shorter when it is colder.

Several exact copies of this bar have been made and are kept securely in different places. There is consequently very little danger of all the bars being destroyed or lost at the same time. All yard measures should be the same length as the distance between the marks on the standard. The yard is divided into three equal parts, and each of these is called a *foot*. A foot is divided into twelve equal parts, and each part is called an *inch*.

**The metric standard of length.**—Lengths are not measured in yards, feet, and inches in all countries. In France, and many other countries, the standard length is what is called a *metre*. "The metre is the length, at the temperature of 0° C., of the iridio-platinum bar, numbered 16, deposited with the Board of Trade." The metre is longer than the yard. You know there are thirty-six inches in the yard, but the metre measures about thirty-nine and one-third inches, or three feet, three and one-third inches (3 feet 3½ inches). (This number is easily remembered because it only contains the figure 3.)

**Divisions of the metre.**—The metre is not divided in the same way as the yard. A much better plan is adopted. In the metric system all the measures are connected with one another by tens, or tenths. The metre is divided into ten equal parts called *decimetres*, each of which is divided into ten equal parts called *centimetres*, and these, in turn, are divided into ten equal parts called *millimetres*.

For lengths greater than a metre the same simple plan is used. A length which contains exactly ten metres is called a *dekametre*; one which just contains a hundred metres is called a *hektometre*; and one which is exactly a thousand times as long as a metre is called a *kilometre*.

All that it is necessary to do in order not to confuse these

<sup>1</sup> This temperature will be understood after studying Chapter IV.

names with one another is to learn the meaning of the prefixes, namely :—

<i>milli-</i>	signifies	thousandth	( $\frac{1}{1000}$ )
<i>centi-</i>	„	hundredth	( $\frac{1}{100}$ )
<i>deci-</i>	„	tenth	( $\frac{1}{10}$ )
<i>deka-</i>	„	ten times	(10)
<i>hekto-</i>	„	hundred times	(100)
<i>kilo-</i>	„	thousand times	(1000)

You will find later on that the same prefixes are used in connection with metric measures of area, volume, capacity, and weight, and as they always signify the same multiples or fractions of the unit, the student who knows their meaning has the key to the whole metric system of weights and measures. The table of Long Measure in this system can now be understood :—

#### METRIC LONG MEASURE.

*The standard is 1 metre, which is equal in length to about  $1\frac{1}{11}$  English yards.*

$\frac{1}{1000}$ metre is	1 <i>milli-</i> metre
$\frac{1}{100}$ „	1 <i>centi-</i> metre or 10 millimetres
$\frac{1}{10}$ „	1 <i>deci-</i> metre, or 10 centimetres, or 100 millimetres
10 metres are	1 <i>deka-</i> metre
100 „	1 <i>hekto-</i> metre
1000 „	1 <i>kilo-</i> metre

**Fractions of units.**—From the foregoing description it has been seen that in the metric system of measurement there is a regular series of tens from the lowest to the highest measure. In the British system, there is no simple connection of this kind. A foot is one-third of a yard, an inch is one-twelfth of a foot, or the thirty-sixth part of a yard, a pole is five and a half yards, and so on.

For smaller lengths than an inch we use such fractions as  $\frac{3}{4}$  inch,  $\frac{1}{2}$  inch,  $\frac{1}{4}$  inch,  $\frac{1}{8}$  inch, and  $\frac{1}{16}$  inch. In the metric system all these troublesome numbers are avoided, and decimals are used instead of vulgar fractions. The units are separated from the parts of a unit by a dot ; tenths are represented by the first number to the right of the dot, hundredths by the second number to the right, thousandths by the third number, and so on. On this *decimal* system of notation, therefore, the number 9376.542 metres means nine thousand, three hundred and

seventy-six metres *plus* five-tenths of a metre, four-hundredths of a metre, and two-thousandths of a metre. And as all the multiples and fractions of the metre are connected with one another by tens, the same number can be expressed as 9 kilometres, 3 hektometres, 7 dekametres, 6 metres, 5 decimetres, 4 centimetres, and 2 millimetres.

## MEASUREMENT OF AREA

**8. Square inches.**—Mark off lengths an inch apart from one corner along the bottom edge of a page of your exercise book or of a sheet of paper. Draw lines straight up the page from each of these marks perpendicular to the edge of the paper. From the same corner mark off lengths of inches on the side edge of your paper, and draw lines across the paper parallel to the bottom edge. You will thus divide the paper into square inches. Count how many square inches there are in one row, neglecting the parts of a square inch which may happen to be over at the end, and then count the number of rows. By multiplying these two numbers together, you will evidently obtain the total number of square inches.

**9. Length and width.**—Draw an oblong 4 inches in length, and 3 inches in width, or height, as it is called in geometry, and divide it into square inches as before. Notice that the number of square inches is equal to the length multiplied by the width.

**10. Rule for areas of oblongs.**—The preceding exercises have shown you that, to find the area of a square or an oblong, it is only necessary to multiply the length of the bottom or base by the height. Use this knowledge to find the num-

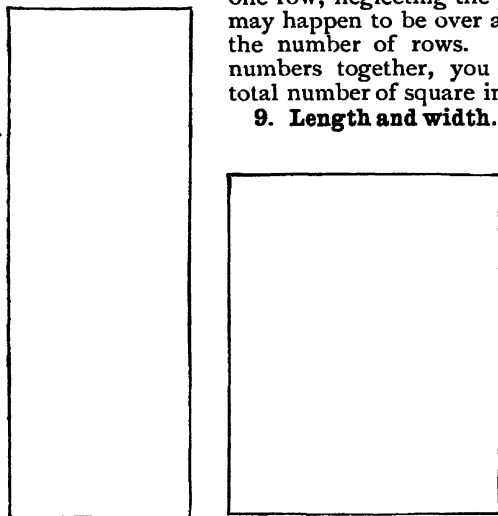


FIG. 3.—How many square centimetres are there in each of these rectangles?

ber of square centimetres in the oblongs drawn in Fig. 3.

These two oblongs could be considered to represent two lengths of materials of different widths.

Multiplication of decimals is performed in the same way as ordinary multiplication. When the sum has been done, count the number of

figures to the right of the decimal point in the numbers multiplied together, and place a decimal point in the answer so that the same number of figures are to the right of it.

**11. Area of square in metric measure.**—Draw a square decimetre and divide it into square centimetres, as shown on a reduced scale in Fig. 4. Count the number of square centimetres.

Notice that you obtain the same result by counting the squares in one row and multiplying the result by the number of rows.

**12. Square inch and square centimetre.**—Draw a square inch and divide it up into square centimetres as shown in Fig. 5. Notice that it takes about six and a quarter square centimetres to make one square inch.

**13. Square foot and square decimetre.**—Draw a square foot upon a sheet of paper, and divide it up into square decimetres. Notice that one square foot is a little larger in area than nine square decimetres.

**14. Applications of measurements of area.**—(a) *Linoleum.*—Measure the length and width of a passage or hall, and find approximately the number of square yards of linoleum required to cover it.

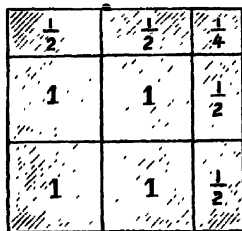


FIG. 5.—Square inch divided into square centimetres.

10									
9									
8									
7									
6									
5									
4									
3									
2									
1	2	3	4	5	6	7	8	9	10

FIG. 4.—Number of square centimetres in one square decimetre. One-quarter actual area.

Linoleums as a rule are made 6 feet wide, and borders can be obtained in widths of 9 in., 12 in., or 18 in. Measure the area of the floor of a square or an oblong room, and find how many yards length of linoleum of the width mentioned would be required to cover it, with or without a border.

(b) *Wall papers.*—English wall papers are sold in pieces 12 yards long and 21 inches wide; or, expressed in feet, 36 feet long and 1 $\frac{3}{4}$  feet wide. Measure the distance around a room, including doors, windows, &c., and the height from skirting to cornice. Find

how many pieces of paper would have to be ordered to paper the walls. The paper not used for doors, &c., may be regarded as waste.

**EXAMPLE.**—Distance around the room, 60 feet ; height from skirting to cornice, 9 feet.

Number of square feet,  $60 \times 9 = 540$ .

Number of square feet in one piece of wall paper,  $36 \times 1\frac{3}{4} = 63$ .

Therefore—

Number of pieces required =  $\frac{540}{63} = 8\frac{1}{2}$  (about).

As the pieces of paper are only sold whole, it would be necessary to order 9 of them to paper the room.

(c) *Carpets.*—Suppose a room in your house is to be carpeted. Find how many yards of a material you would require to use as a border or “surround,” and supposing that this is 18 inches wide, what size carpet would be required to cover the remainder of the floor.

Find how many yards of Brussels carpet, 27 inches wide, would be required to cover the floor of a room you can conveniently measure. You must, of course, remember that all the carpet you buy will be 27 inches wide, so that possibly there may be a waste strip.

**Measurement of area.**—Cloth, or other dress material, is sold by the yard length, but every one knows that the actual amount of material obtained depends also upon the width. Silk and satin have usually a width of from 18 to 22 inches, and from 14 to 16 yards are required to make a plain costume. Velvet is generally made slightly wider than silk or satin, so a shorter length—about 10 or 12 yards—is sufficient for a costume. Double width tweeds are from 54 to 56 inches wide, so only 4 or 5 yards length are required ; that is, less than one-third the length necessary when single-width silk is used.

Many other familiar instances of a similar kind continually occur in connection with domestic affairs. If the floor of a room has to be carpeted, you know very well it would not be enough to measure the length of the room only, nor its width only, because both of these are measures of length. To know how much carpet is necessary, the amount of surface the floor has, or what is called its *area* must be found. To do this, both the length and width of the floor must be measured, and if the room is a square or an oblong one the area can be found by multiplying the two numbers together. If the length and width are measured in feet, then, by multiplying them together, the area of the floor in square feet is obtained ; if the measurement of the length and width are taken in inches, the area in square inches is obtained by multiplying the two numbers together.

Whenever areas are measured in this country, square inches, square feet, square miles, or some other unit from square

measure is employed. "Square measure" is obtained from "long measure" by multiplying. Thus, as there are 12 inches in a foot, there are  $12 \times 12$  square inches in a square foot. This will be understood by examining Fig. 6. Each of the large squares bounded by thick lines represents a square foot, but it is of course smaller than a real square foot.

**British square measure.**—By referring to Fig. 6, which illustrates how a square yard may be divided into square feet

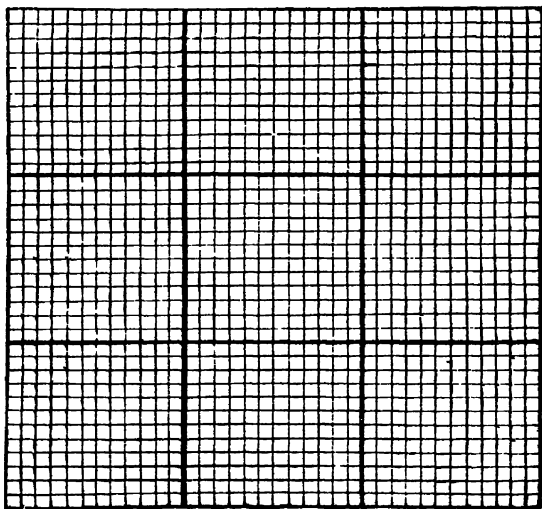


FIG. 6.—This represents a square yard on a small scale. Each small square represents a square inch, and the large squares bounded by thick lines represent square feet.

and square inches, it is easy to see how square measure is obtained from long measure. This may be put in the form of a table, thus :

144	(= 12 × 12)	square inches	make	1	square foot.
9	(= 3 × 3)	„ feet	„	1	„ yard.
$30\frac{1}{4}$	(= $5\frac{1}{2} \times 5\frac{1}{2}$ )	„ yards	„	1	„ pole.

How many square inches are there in a square yard? You can find out by counting the small squares in Fig. 6, or by counting the squares in one of the areas representing a foot and multiplying the number by nine. •



**Metric square measure.**—If instead of measuring in feet the length and breadth of a floor to be covered with carpet the measurements are made in metres or decimetres, it is evident that the area obtained by multiplying the numbers together would not be in square feet, but in what are called square metres, or square decimetres. Square measure in the metric system is obtained from metric long measure in a similar way to that already described, for British square measure. All we mean by the metric system is the plan of using metres, etc., instead of yards, etc., in measurements of all kinds. We can now write down the measures of area or surface in the metric system :

100 (= 10 × 10) sq. millimetres make 1 sq. centimetre.

100 (= 10 × 10) „ centimetres „ 1 „ decimetre.

100 (= 10 × 10) „ decimetres „ 1 „ metre.

A square decimetre is too large to be shown on this page. Fig. 7 shows a row of ten square centimetres, but there are ten rows of this kind in a square decimetre. One of the centimetres is divided into square millimetres. As 1 square centimetre contains 100

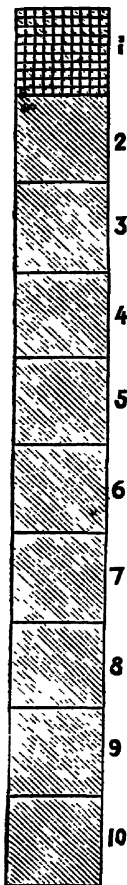


FIG. 7.—Square centimetres and square millimetres. Ten rows of ten square centimetres make one square decimetre.

square millimetres, 1 square decimetre or 100 square centimetres contains 100 × 100 square millimetres, that is, 10,000 square millimetres.

The size of a square centimetre is compared with the size of a square inch in Fig. 8. You will see that a square inch is much larger than a square centimetre. Each side of the square inch is 2·54 centimetres in length, so the number of square centimetres

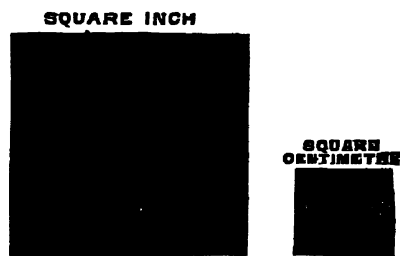


FIG. 8.—Comparative sizes of square inch and square centimetre.

in a square inch is  $2.54 \times 2.54 = 6.45$ , or nearly  $6\frac{1}{2}$  square centimetres.

The relation between a square foot and a square decimetre can be considered in a similar way. Thus, each side of a square foot is a little more than three decimetres long. The area of one square foot, expressed in decimetres, is therefore a little more than  $3 \times 3 = 9$  square decimetres. The exact number of square decimetres in one square foot is  $9.29$ , that is, about  $9\frac{1}{3}$ .

### SUMMARY.

**Measurement of length.**—In estimating lengths, a standard, or unit, of length is necessary.

The *British standard of length* is the *Imperial standard yard*. It may be defined as “the length, at  $62^{\circ}$  F., marked on a bronze bar deposited with the Board of Trade.”

The yard is *divided* into three equal parts, each called a *foot*. A foot is divided into twelve equal parts, each called an *inch*.

The *Metric standard of length* is called a *metre*. In the metric system all measures are connected with one another by *tens* or *tenths*.

The metre is divided into equal parts called *decimetres*, each of which is again divided into ten equal parts called *centimetres*, and these in turn are divided into ten equal parts called *millimetres*.

A length which contains exactly 10 metres is called a *dekametre*, one which just contains 10 dekametres, or 100 metres, is called a *hektometre*, and one which is exactly 1000 times as long as a metre is called a *kilometre*.

**Measurement of area.**—To estimate area, or surface, it is necessary to measure in two directions, *i.e.*, to find the length and breadth of the surface.

By multiplying together the length and breadth of a square or an oblong we obtain its area.

**British square measure.**—Measures of area are called *square measures*. In British square measure, 144 square inches make one square foot; 9 square feet make one square yard; and  $30\frac{1}{4}$  square yards make one square pole.

**Metric square measure.**—100 square millimetres make one square centimetre; 100 square centimetres make one square decimetre; and 100 square decimetres make one square metre.

### EXERCISES ON CHAPTER I.

1. What is meant by the Imperial standard yard? Name the parts into which it is divided.
2. What is a metre? Compare its length with that of a yard.
3. Explain how the metre is divided, and write down the names which are given to these parts.
4. Which length is the greater,  $1\frac{1}{8}$  inch, or  $1.3$  inch?

5. A certain length was measured and found to be  $25\frac{1}{2}$  inches. Express this length in feet, inches, and decimals of an inch.

6. A window 3 feet wide has to be fitted with a brass rod for curtains. What length of rod would be required, supposing that  $\frac{1}{2}$  inch has to be allowed at each end of the rod so that it shall not be too tight?

7. Add  $\frac{1}{8}$  inch to each of the following lengths:  $8\frac{1}{8}$  inches,  $15\frac{1}{4}$  inches,  $13\frac{3}{4}$  inches,  $11\frac{1}{2}$  inches,  $6\frac{3}{8}$  inches.

8. What is the area of a piece of Tussock silk 9 yards long and 22 inches wide?

9. Suppose a material sold in a width of 44 inches has to be backed with a material 22 inches wide. If 7 yards of the merino are required for a costume, how many yards of the lining material would have exactly the same area?

10. What length of carpet border would be required to go round a room 12 feet wide and 12 feet long? If the border is 18 inches wide, what would be the length and width of the centre of the room not covered by it?

11. What area of linoleum would be required to cover a passage 5 feet wide and 18 feet long? Suppose the linoleum is 6 feet wide, what length would you order, assuming the pattern will fit any way?

12. French wall papers are made 8 metres long and 45 centimetres wide. Express these lengths approximately in English measure.

13. How many pieces of wall paper 12 yards long and 21 inches wide would be required to paper a room 9 feet high from skirting to cornice, and 60 feet round? Do not make any allowance for doors, windows, &c.

## CHAPTER II

### CUBIC AND FLUID MEASURE

#### EXERCISES IN MEASUREMENT OF VOLUME AND CAPACITY

**15. Cubic inch.**—Obtain a thick piece of soap, and cut one end so that it is square with two sides, as indicated in Fig. 9. From one corner, where the end and these two faces meet, mark off the length of an inch

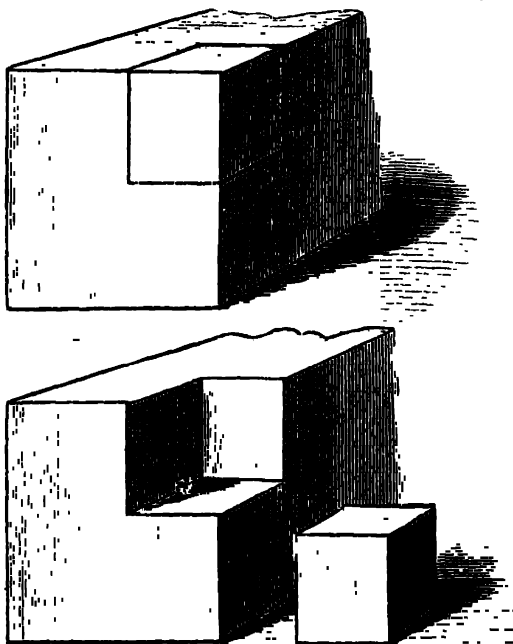


FIG. 9.—How to mark and cut out a cubic inch or cubic centimetre from a piece of soap.

along the three edges. Scratch a line across the soap from each of the points thus obtained, and square with the edge. You will then have three square inches marked upon the soap. By cutting deep along

these lines a cube of soap having each of its edges an inch in length, and each of its faces a square inch in area, is obtained. This is a cubic inch.

To test whether your cubic inch is true it could be compared with one of the wooden cubes sold for Kindergarten teaching.

**16. Cubic centimetre.**—Using soap, cheese, clay, or similar material as before, cut one end square with the faces, and then mark off a length of one centimetre along three edges meeting at one corner. Scratch

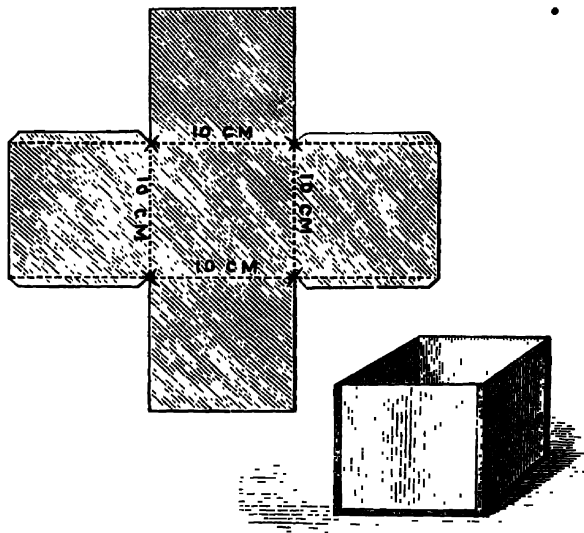


FIG. 10.—Construction of a box to hold 1 cubic decimetre, that is, 1 litre.

lines from each of these points so as to mark three square centimetres upon the material used, and then cut along the lines to obtain one cubic centimetre, as in the previous exercise. Compare this cube with the cubic inch. If you have time and material, cut a rod of plastic substance sixteen centimetres long and one square centimetre at each end. Divide this into sixteen cubic centimetres, and then show that the cube made by pressing these together is of about the same size as a cubic inch.

**17. Standard capacity boxes.**—(a) Cut out of cardboard a figure of the shape shown in Fig. 10, and of the size indicated by the numbers. Cut the cardboard half-way through at the dotted lines, and then bend it to form a cubical box. Bind the edges together with glued tape, and varnish the box inside and out to make it watertight. The box will hold 1000 cubic centimetres of a liquid, and this amount is called a *litre*.

(b) Make a similar box having sides one inch in length, so that it will

hold one cubic inch. Also, if time permit, make a box to hold one cubic centimetre.

**18. Fluid measure.**—Obtain a half-pint graduated glass measure, such as is used by photographers (Fig. 11). Examine the divisions upon it: they represent fluid ounces, or parts of fluid ounces, in Apothecaries' Measure.

**19. Capacities of spoons.**—Fill a tea-spoon with water, and pour the water into the graduated measure. Repeat the operation eight times. The water will then be near the level of the 1 oz. mark. You will see from this that as eight tea-spoonfuls of liquid are equal

to one fluid ounce, one tea-spoonful of liquid is equal to one-eighth of a fluid ounce, that is, one drachm. Repeat the exercise, using a dessert-spoon, and then a table-spoon, instead of the tea-spoon. You will find that a dessert-spoon holds about one-quarter of a fluid ounce, and a table-spoon about half an ounce of a liquid.

**20. Drops.**—Obtain a medicine dropper, or a short piece of narrow glass tubing contracted slightly at one end (Fig. 13). Suck water into it, and then let the water drip by drops

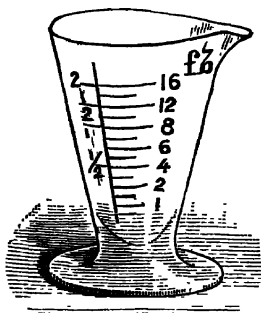


FIG. 12.—Measuring glass graduated into drachms.

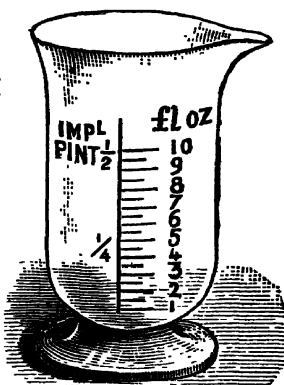


FIG. 11.—Measuring glass graduated into fluid ounces.

into a tea-spoon. Count the number of drops required to fill the spoon. It will be found to be about sixty, showing that one drop of water is roughly equal to one minim.

**21. Imperial pint.**—Obtain an Imperial pint bottle. Find the number of fluid ounces it will hold. If possible, test with your glass measure whether one or two so-called pint or half-pint jugs, or tumblers, are really true measures.

**22. Pint and litre.**—Compare an Imperial pint bottle with a litre flask, such as Continental wines are frequently sold in. Using your measuring glass, show that the number of fluid ounces in one litre is about 35; so that one litre is equal to  $1\frac{1}{2}$  pints.

**23. Graduated measuring jar.**—Obtain a jar graduated into cubic centimetres, like that in Fig. 19. Test the accuracy of your cubic

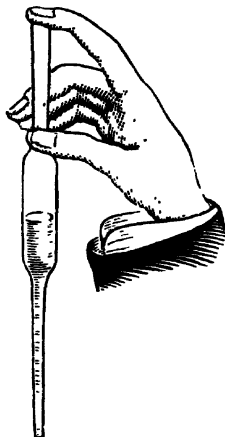


FIG. 13.—Pipette, or medicine dropper.

decimetre box by pouring 1000 cubic centimetres of water, as measured by the graduated jar, into the box, and noticing whether the water just fills the box.

Fill a measuring jar about half way up with water. Notice the level. Hold a cubic inch of wood in the water by means of a long pin (Fig. 14). Notice the rise of level. The difference between the two levels evidently represents the number of cubic centimetres equal to one cubic inch.

**24. Burette.**—Examine a burette (Fig. 19). Notice that the divisions are numbered from the top downwards. Fix the burette upright in a clamp, and about half fill it with water. Notice the level. Find the volume of a stick of slate pencil by gently lowering it into the water and noticing the rise of level produced. (The pencil must, of course, be completely immersed in water.)

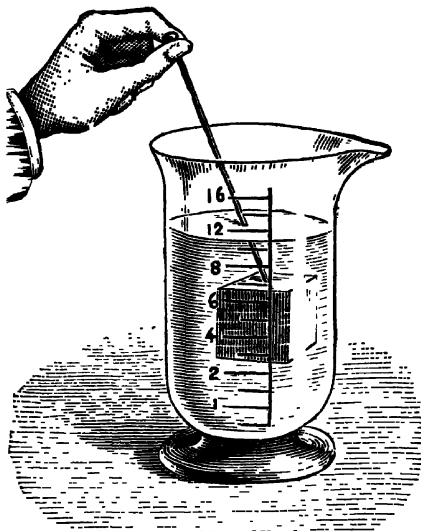


FIG. 14.—A method of finding the number of fluid ounces or cubic centimetres equal to one cubic inch.

**Cubic measure.**—The size of a solid, or the amount of room it takes up, or the space it occupies, is called its *volume*.

The volume of a solid body is obtained by measuring in three directions. Just as to find the area of a surface it is necessary to measure length and breadth, so to measure the volume of a solid we must find, in addition to measurements of length and breadth, another distance called the thickness. If length, breadth, and thickness are multiplied together in

the case of a regular solid, the result obtained represents the volume or cubical content.

An examination of Fig. 15 will assist the student to understand this result. The single small cube at the extreme right may be considered to represent one cubic inch. Next to it is a row of twelve such cubes, and then a slab made up of twelve of these rows on top of one another. Finally, there is the large cube made up of twelve slabs placed face to face. The large cube, which represents a cubic foot, may thus be regarded as made up of 12 slices, each containing 12 rows of 12 cubic inches.

The total number of cubic inches in 1 cubic foot is therefore  $12 \times 12 \times 12$ , that is, 1728. Or, to put it another way, the length of each edge of the large cube is 12 inches, the area of each face is  $12 \times 12$  square inches, that is, 144 square inches, and the volume is equal to the area of a face multiplied by the thickness, that is,  $144 \times 12$ , or 1728 cubic inches.

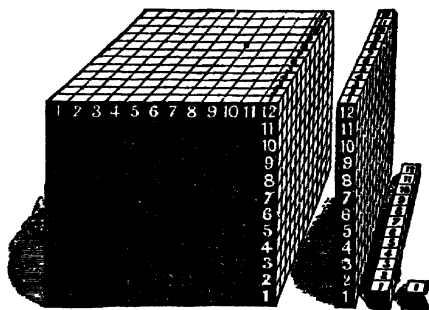


FIG. 15.—To explain why there are 1728 cubic inches in one cubic foot.

The number of cubic feet in a cubic yard can, in a similar way, be found to be equal to  $3 \times 3 \times 3$ , that is, 27 (Fig. 16).

**British measures of capacity.**—The amount of fluid a

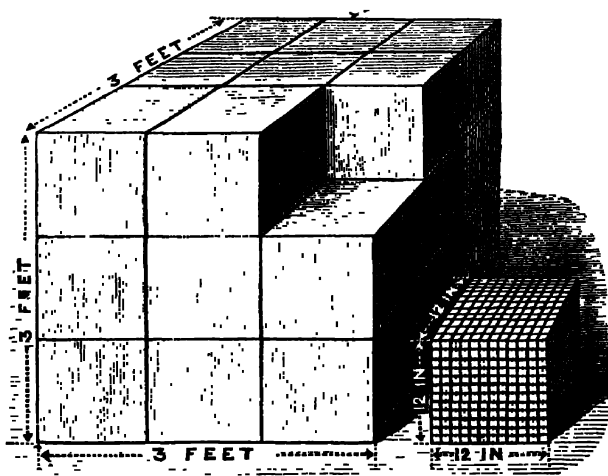


FIG. 16.—Relation of cubic foot to cubic yard.

vessel will hold is known as the capacity of the vessel. Familiar British measures of capacity are the pint, quart, and gallon.

4 gills = 1 pint.  
2 pints = 1 quart.  
4 quarts = 1 gallon.



This is known as the Wine and Spirit Measure, but with the exception of the gill the same measure is used for all liquids. The gill is also called a noggin or quartern.

For measuring fluids for medicines, what is known as Apothecaries' Measure is used.

- 1 minim = 1 drop of thin liquid.
- 60 minims = 1 drachm (about a tea-spoonful).
- 8 drachms = 1 fluid ounce (4 dessert-spoonfuls or 2 table-spoonfuls).
- 20 fluid ounces = 1 pint.

The Imperial gallon has a volume equal to  $277\frac{1}{4}$  cubic inches. An Imperial pint is one-eighth of a gallon, and therefore holds about  $34\frac{3}{8}$  cubic inches. This is true for Wine and Spirit Measure, Beer Measure, or Apothecaries' Measure, but only in the last case is the pint divided into 20 fluid ounces, which are again divided into drachms and minims for the measurement of the small quantities required in making up physicians' prescriptions. A pint of milk should, of course, be the same amount as a pint of beer or any other liquid, but beers are often sold in bottles which are not intended to be pints, and are known merely as bottles of beer or reputed pints. If, however, a bottle is marked an Imperial pint, it should contain 20 fluid ounces of liquid.

Screw-stoppered bottles usually hold Imperial pints, but Guinness's stout and Allsopp's or Bass's pale ale are sold in reputed pints and quarts as well as Imperial pints.

**Metric measures of volume.**—As all metric measures are connected with one another by tens, it is easy to deduce metric measures of volume from those of area or length. The small cube, A, at the left in Fig. 17, may be regarded as representing 1 cubic centimetre; 10 such cubic centimetres make the row, B, shown; 10 rows of 10 cubic centimetres make the slice, C; and 10 such slices make up a complete cube, D. The number of cubic centimetres in the large cube is therefore 10 slices of 10 rows of 10 cubic centimetres, that is, 1000 cubic centimetres or 1 cubic decimetre.

The subject can be looked at in another way. You know there are 10 centimetres in a decimetre, so you may say the

edge of the decimetre cube is 10 centimetres in length ; the area of one of its faces is  $10 \times 10 = 100$  square centimetres ; and its volume is  $10 \times 10 \times 10 = 100 \times 10 = 1000$  cubic centimetres.

Sixty-one cubic inches are about equal to the volume of a

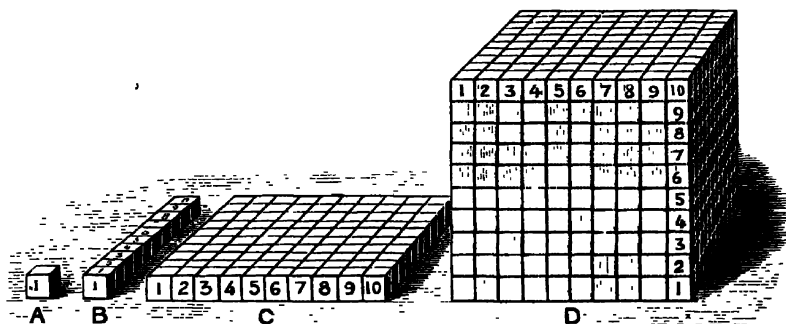


FIG. 17.—To show the relation between 1 cubic centimetre and 1 cubic decimetre.

cubic decimetre ; and one cubic inch is equal roughly to  $16\frac{1}{2}$  cubic centimetres (Fig. 18).

**The litre.**—If a hollow cube is made 1 decimetre long, 1 decimetre broad, and 1 decimetre deep, it will hold 1000 cubic

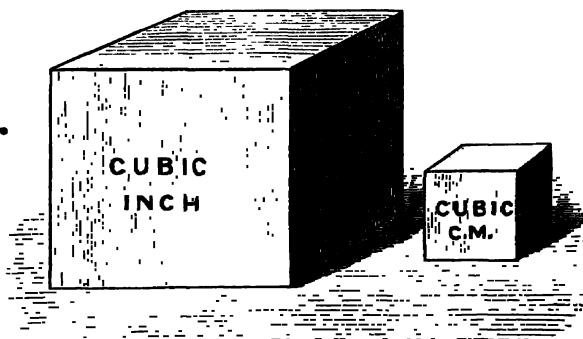


FIG. 18.—Cubic inch and cubic centimetre in their true sizes.

centimetres of liquid. This capacity is called a *litre*. All liquids are measured in litres in countries where the metric system is adopted. Thus, in France, wine, milk, and such liquids, are sold by litres and half-litres, instead of by quarts

## EXPERIMENTAL HYGIENE

and pints. A litre is equal to about one and three-quarters English pints.

If a measuring glass, capable of holding 1 litre, has its capacity divided into 1000 equal parts, then the capacity from one division to the next represents 1 cubic centimetre. Suppose such a glass is half filled with liquid, and some small objects are dropped into it until the level of the liquid rises 200 divisions. Then without any calculation you know that the volume of the objects is 200 cubic centimetres. It is often convenient to determine the volume of a thing in this way with a metric measuring glass.

The fractions and multiples of a litre are named upon precisely the same plan as those of the metre already described.

### METRIC MEASURES OF CAPACITY.

*The standard is 1 litre, which is equal to 1000 cubic centimetres, or about  $1\frac{3}{4}$  English pints.*

1000	litre is	1 milli-litre.
100	„ „	1 centi-litre.
10	„ „	1 deci-litre.
10	litres are	1 deka-litre.
100	„ „	1 hekto-litre.
1000	„ „	1 kilo-litre.

**Graduated measuring vessels.**—It has been seen that the standard of capacity in the metric system of measurement is the litre. Since in scientific work the metric units of capacity are made use of in all countries, it is customary to employ measuring vessels, graduated metrically, of different shapes according to the purpose for which they are intended. Each of these vessels contains so many litres, or in the case of the smallest ones a certain fraction of a litre.

The vessels have different names, according to their shapes. The illustration shows the commonest shapes in which measuring vessels are made. The meaning of each of the names *flask*, *jar*, *burette*, and *pipette* will be at once understood from Fig. 19.

The way in which each of these things is used will be best learnt gradually by experience. In the different experiments, which the student will, it is to be hoped, perform for himself, the variously shaped vessels are often used.

Since a litre contains a thousand cubic centimetres, it is clear the flask in the illustration labelled 500 c.c. contains half a litre.

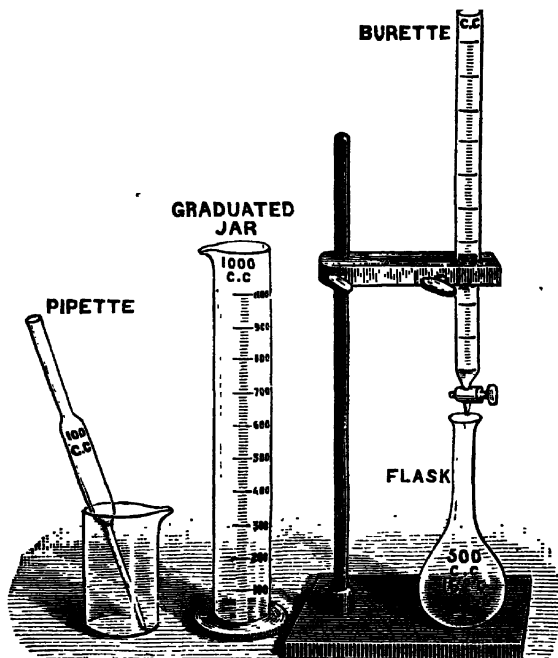


FIG. 19.—Graduated measuring vessels.

Similarly up to the 1000 c.c. mark in the graduated jar will be one litre. The burette holds over one-tenth of a litre, and so on.

### SUMMARY.

**Cubic measure.**—The *volume* of a solid is its *size*, or the *amount of room* it takes up, or the *space it occupies*.

Volume is obtained by multiplying together the length, breadth, and thickness of a rectangular solid, the measurements being at right angles to one another.

**British cubic measure.**—1728 cubic inches make one cubic foot, and 27 cubic feet make one cubic yard.

**Metric cubic measure.**—1000 cubic centimetres make one cubic decimetre, 1000 cubic decimetres make one cubic metre.

The *litre* is the metric measure of *capacity*. It is the volume of

liquid which a hollow cube of a decimetre edge will hold, and is equal to 1000 cubic centimetres.

**British measures of capacity.**—Familiar British measures of capacity are the *pint*, *quart*, and *gallon*. 4 gills make one pint, 2 pints make one quart, and 4 quarts make one gallon.

*Apothecaries' measures of capacity.*—A drop of thin liquid has a capacity of about one *minim*, 60 minims make one *drachm*, 8 drachms make one *fluid ounce*, and 20 fluid ounces make one *pint*.

**Determination of volume of irregular solids.**—The volume of a solid of any shape can be measured by the amount of a liquid it will displace when plunged into it. If the solid is lighter than the liquid, it can be held beneath it by means of a hat pin.

#### EXERCISES ON CHAPTER II.

1. One pint is equal to 34.7 cubic inches, and 1 inch is equal to 2.54 centimetres. How many pints are there in 1000 cubic centimetres?
2. How would you determine the volume of a pebble in cubic centimetres?
3. How many litres are there in a gallon of water?
4. How many glasses of wine, each holding  $\frac{1}{4}$  pint, could you obtain from a litre bottle of wine?
5. Which is the greater, a quarter of a litre or half a pint?
6. Describe how you would cut out a cubic inch of soap. How could you demonstrate that it takes about 16 cubic centimetres to make 1 cubic inch?
7. If you had a British measuring glass and wished to test whether a bottle held an Imperial pint or not, how would you proceed?

## CHAPTER III

### WEIGHT AND DENSITY

#### WEIGHTS AND WEIGHING

**25. The balance.**---(a) Uncover the balance and identify the different parts by reference to Fig. 20. Raise the beam, AB, of the balance, off the supports by turning the handle, C. Notice whether the pointer, F, swings equally on both sides of the middle of the scale, G : if it does, the balance is ready for use ; but if not, let down the beam and turn the

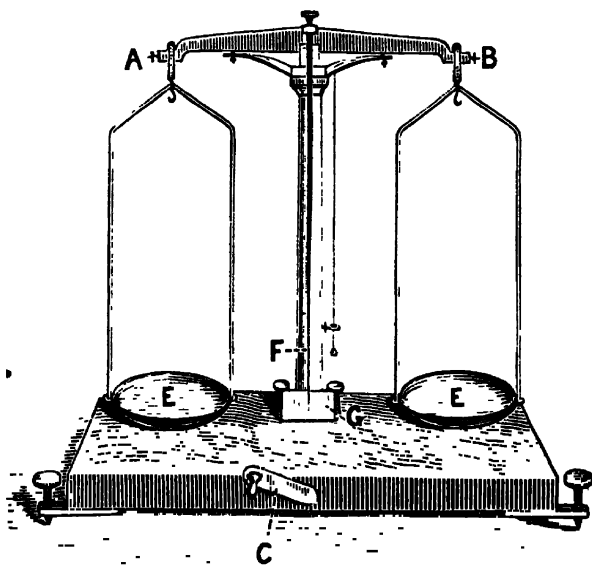


FIG. 20.—A Students' Balance.

small screw at B, then try again. Repeat this adjustment until the swings to right and left are equal.

**26. The method of weighing.**—Put some shot in a watch-glass, and place it in the *left-hand pan*, E, of the balance. Examine your box of weights, and, selecting a weight which you estimate to be about the same as the shot, take hold of it *with the forceps* and place it in the *right-hand pan* (the weights should always be placed in this pan).

Now slightly raise the beam to see whether the estimated weight is nearly equal to the weight of the shot. If the weight appears a little below what is wanted, pick up, with the forceps, the next heaviest weight in the box, and try it in the pan with the other. If the two together are too heavy, take out the smaller weight and put in the one below it, and so on, adding one weight after the other, *without missing any*, until you find the correct weight. When you have completed the weighing, write down and add up the weights that are missing from their places, and check the figures as you put each weight back in its place.

**27. Drams and ounces.**—Place an ounce weight in one pan of a balance, and cut off a piece of metal wire weighing a little more than 1 ounce. Reduce the length of the wire gradually by cutting, until it

weighs exactly 1 ounce; then cut the wire into 16 equal pieces. Each of the pieces will weigh 1 dram.

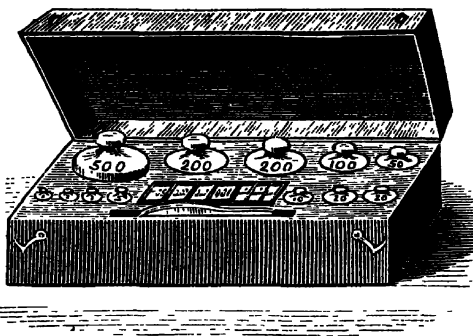


FIG. 21.—Box of metric weights.

**28. Fluid ounce and an ounce weight.**—Place

a measuring glass in one pan of a pair of scales, and balance it with lead or nails put into the other pan. Pour a *fluid* ounce of cold water into the measuring glass. Notice that an ounce weight will exactly balance it. A fluid ounce of water is thus shown to weigh an ounce.

**29. Metric weights.**—

Examine a box of metric weights. The numbers upon the weights represent grams and fractions of grams. Write a list of all the weights having numbers upon them.

**30. Grams and ounces.**—Place a weight of 100 grams in one pan of a balance, and show that it can be balanced by  $3\frac{1}{2}$  ounces in the other pan.

**31. Kilogram and pound.**—Put a kilogram (1000 grams) in one pan of a *strong* balance, and show that it can be balanced by a little less than 2 pounds  $3\frac{1}{2}$  ounces.

**32. Exercises in weighing.**—Find the weight in ounces and parts of an ounce of each of the following: a table-spoonful of flour, and a table-spoonful of best brown sugar. To do this, first find the weight of a watch-glass, or other suitable vessel to hold the powder, then transfer the powder to be weighed from the spoon to the watch-glass. Find the weight of both together, and after subtracting the weight of the watch-glass, the weight of the powder will be obtained.

In a similar manner find the weight, in succession, of a teacupful of

flour, ground coffee, powdered loaf sugar. First find the weight of the empty cup, and then when it is full of one of the substances. When the weight of the cup has been subtracted from the total weight, the remainder will give the weight of the flour, coffee, &c.

**The need for weighing.**—It is very often necessary in the household, sometimes in the kitchen, sometimes in the sick room, and on other occasions, to be able to measure the amount of different materials to be employed for various purposes. In the kitchen, for instance, this fact is recognised, and the cook has a number of rough rules she employs. But these are apt to lead to error because of the uncertainty about the standards adopted. For instance, one rule of the kitchen says that a lump of soft butter the size of an egg weighs an ounce. But it is not difficult to understand that butter varies very much in softness, and eggs vary much in size. So that sometimes one amount of butter will be used and sometimes another. It is not very surprising that the results of different operations vary considerably when such rough-and-ready rules are in use. Another kitchen rule is, that a level table-spoonful of flour weighs half-an-ounce. But it depends upon how much pressing the cook does as to how much flour she can get into the table-spoon. All such inaccuracy is avoided when everything is weighed. An ounce of butter is a certain definite amount that can always be repeated. A pound of sugar means a certain amount that is unaffected by the size of any variable thing. To be sure of good results then, the cook should make a habit of weighing everything.

**The principle of the balance.**—A simple form of balance, or pair of scales, is represented in Fig. 22. The beam, AB, is supported at the centre, C, and if the balance is true, the scale pans have no tendency to dip either on one side or the other. When, as in this case, the pans are suspended from points at equal distances from the pivot, any weight in one pan is balanced by an equal weight in the other. In the case of every balance, if the beam swings evenly on both sides of the pivot, the following principle of equality holds good :—

$$\begin{array}{l} \text{Weight on} \\ \text{one side} \end{array} \times \begin{array}{l} \text{Distance} \\ \text{from pivot} \end{array} = \begin{array}{l} \text{Weight on} \\ \text{other side} \end{array} \times \begin{array}{l} \text{Distance} \\ \text{from pivot.} \end{array}$$



This kind of balance is good enough for ordinary purposes, but when exact weighings are wanted, a better form, such as that shown in Fig. 20, is used. All the parts in this balance are very carefully made, and the greatest possible pains are taken to have very delicate supports and accurate adjustments. A brass beam is supported at its middle line on a knife edge of hard steel, which, when the balance is in use, rests on a true surface of similar steel. The hooks to which the pans are attached are likewise provided with V-shaped grooves of hard steel (Fig. 20), which, when the balance is in use, rests upon knife edges on the upper parts of the beam. To the middle of the beam is attached

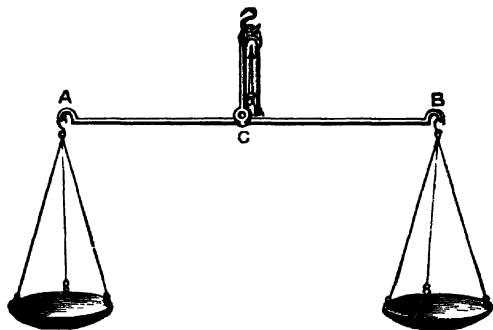


FIG. 22.—A simple balance.

a pointer, the end of which moves in front of an ivory scale, fixed at the bottom of the upright which carries the beam. When not in use, the beam and hooks are lifted off the knife edges by turning a handle.

**Measurement of weight.**—Just as in measuring lengths we found it was necessary to have a standard with which to compare them, so in measuring weight we must also have a standard or unit. Then we can say how many times the weight of a given body is greater or smaller than our unit. In this country the standard of weight is that “of a piece of platinum weighed in vacuo at the temperature of  $0^{\circ}$  C., which is deposited with the Board of Trade.” This lump of platinum is called the *Imperial standard pound avoirdupois*. The divisions, etc., of the Imperial pound you have already learnt at school under the name of “avoirdupois weight.”

A weight of 1 lb. avoirdupois is kept at a Standards Office in every city, so as to test the lb. weights used by tradesmen, and see whether they really weigh 1 lb. or are too light.

*Avoirdupois Weight.*

16 drams	make	ounce ( $437\frac{1}{2}$ grains).
16 ounces	„	pound (7000 grains).
14 pounds	„	stone.
28 pounds	„	quarter.
4 quarters	„	hundredweight.
20 hundredweights	„	ton.

**The kilogram and gram.**—The standard of weight which is adopted in France, and in other countries where they use the metric system, is called the *kilogram*. The kilogram is bigger than the British pound; indeed it is equal to about two and one-fifth of these pounds. It is interesting to know how the weight of a kilogram is obtained. It is agreed to give the name *gram* to the weight of water which a little vessel holding one cubic centimetre contains. (To be exact, the temperature of the water must be  $4^{\circ}$  C., but this need not concern the student now.) The kilogram has the same weight as one thousand cubic centimetres of water, that is, as a litre of water. The names used for the divisions, etc., of the gram are obtained in the same way as in the case of the metre, thus:

*Metric Measurement of Weight.*

10 milligrams = 1 centigram.	10 grams = 1 dekagram.
10 centigrams = 1 decigram.	10 dekagrams = 1 hektogram.
10 decigrams = 1 gram.	10 hektograms = 1 kilogram.

**How to remember metric measures.**—As the metric measures of length, volume, and weight have all now been described, it may be helpful to bring them together to show how the fractions and multiples of the units have the same prefixes, namely, milli, centi, deci, deka, hekto, kilo.

By putting these words in front of the words metre, litre, and

gram, the metric measures of length, volume, and weight are obtained, as shown in the following table :—

	Length.	Volume.	Weight.
$\frac{1}{1000}$	Milli-metre	Milli-litre	Milli-gram
$\frac{1}{100}$	Centi-metre	Centi-litre	Centi-gram
$\frac{1}{10}$	Deci-metre	Deci-litre	Deci-gram
1	<b>Metre</b>	<b>Litre</b>	<b>Gram</b>
10	Deka-metre	Deka-litre	Deka-gram
100	Hekto-metre	Hekto-litre	Hekto-gram
1000	Kilo-metre	Kilo-litre	Kilo-gram

You see from this that what you have learnt to call the metric system of weights and measures is much simpler than ours, and the pupils in countries where it is used have not to learn so many different tables as they have in England when they begin "weights and measures" sums.

**Relation of British and metric weights.**<sup>1</sup>—It is a great pity that the simple metric system of weights and measures is not in use in this country. Fortunately, more people each year learn the advantages of the metric system, and, perhaps, some time in the future, the clumsy avoirdupois weights and the complicated measures of length in use in our country will give place to the simple, easy measures in use on the Continent. Until this happens it is worth while to point out some of the ways in which the two systems are related. To begin with measures of weight.—The English grain is the same in every one of our measures of weight. In the apothecaries' ounce there are 480 of these grains, and in the ounce avoirdupois, there are  $437\frac{1}{2}$  of them. The grain is equal to about 0.65 of the decigram ; or, it may be said, the gram contains 15.43 grains. The ounce avoirdupois contains 28.35 grams.

<sup>1</sup> Since the word weight is always used in everyday language whether what is properly called "mass" is meant or not, no attempt has in this book been made to distinguish between "mass" and "weight." If the pupil proceeds later to a formal study of physics, he will have little difficulty in distinguishing between the terms. At this stage the distinction is troublesome and unnecessary.

**Science is based upon measurement.**—It cannot be too soon learnt that science is merely organised common sense. There is a scientific, or right way, and a wrong method of doing everything. The scientific person never trusts to luck and never guesses. The scientific cook, that is the cook who understands her business, when making porridge, for example, does not guess how much oatmeal and how much water are necessary. She weighs the oatmeal and measures the water. In making a cake she carefully weighs each of the ingredients, using exactly the amount she has found, or others have found, by experience, to be necessary to produce a cake of a certain kind. Similarly, such a cook weighs a joint before roasting it, and again after it is cooked, so that she may know just what weight disappears in the process. This desire for accuracy is not “stinginess” nor is it “faddishness”; it is just common sense applied to prevent any anxiety as to results. “There is,” says the wise cook, “a cause for every effect; if I arrange the causes correctly, I need have no fear as to the effects.”

### DENSITY.

**33. Meaning of density.**—Cut a cubic centimetre out of soap, cheese, clay, and cork, in the manner described on p. 14. Weigh each of the cubes in succession, and record their weights in a column as follows:

Weight of cubic centimetre of soap,	.....	grams.
„ „ „	cheese,	..... grams.
„ „ „	clay,	..... grams.
„ „ „	cork,	..... grams.

Notice that the weight of a cubic centimetre differs, thus proving that the weight of the same volume of these different substances is not the same.

**34. Liquids have different densities.**—Scratch a mark on the neck of a small medicine bottle, or stick a piece of gummed paper round the outside of the neck. Weigh the empty bottle when it is clean and dry inside and outside. Now fill the bottle up to the mark with water, and weigh the bottle full of water. Subtract the weight of the bottle, and so obtain the weight of water which fills the bottle up to the mark. In a similar manner find the weight of vinegar, milk, paraffin oil, and other liquids, which fill the bottle up to the mark. Of course the weight of the bottle must be subtracted in each case. Record your results, thus:

Weight of water which fills bottle up to mark,	.....	grams.
„ vinegar „ „ „	.....	grams.
„ milk „ „ „	.....	grams.
„ paraffin oil „ „ „	.....	grams.

Notice that the amounts of different liquids which fill the bottle up to the mark, that is, the same volumes of different liquids, have not the same weight.

**35. The weight of a cubic centimetre of water.**—Place a small beaker (a small wide-mouthed bottle made of thin glass will do) in the left-hand pan of a balance, and counterpoise it with some nails placed in the other pan. Using the burette in the manner described on p. 21, run from it into the beaker 10 c.c. of distilled water. Find what weight in grams must be added to the right-hand pan to counterpoise the beaker again. Similarly, running in 10 c.c. at a time, find the weight of 20, 30, 40 c.c. of water. Determine the weight of a cubic centimetre of water by dividing the weights found by 10, 20, 30, etc.

It will be found that the weight of a cubic centimetre of distilled water is very nearly one gram. It would be exactly a gram if the temperature of the water were 4° C.

**What is meant by density.**—If different solids of the same size or volume are obtained, as every one knows, they may have different weights. If the weight of a cubic centimetre of cork, cheese, soap, clay, is determined, one after the other, with the help of a balance, the clay will be found to be heaviest, the soap will come next, and then will follow the cheese and cork in order. Or, equal volumes of these different solids have different weights.

By filling a bottle up to a mark with different liquids in succession, it can be shown that equal volumes of the liquids used have not the same weight.

Equal weights of iron and water have very different sizes. While a pint of water weighs a pound and a quarter, you know very well that iron weights of a pound and a quarter-pound have not together a volume of anything like a pint.

Two facts are taught by experiments of this kind ; they are:—

1. Equal volumes of different materials may have different weights.

2. Equal weights of different substances may have different volumes.

It is more common to say, however, that materials have different densities. To make this quite clear, consider the familiar example of a pound of feathers and a pound of lead. They have the same weights, but the size, or volume, of the feathers is very much greater than that of the lead. The material in the lead must be packed more closely than in the feathers, which accounts for the lead taking up less room. A

short way of expressing all this is to state that the lead is denser than the feathers.

If the weight of a very small thing is great, it is said to be very dense, or to have a high density. But if an object of large size has a small weight its density is low. Lead has a high density ; feathers have a low density.

**The comparison of densities.**—It is easy to compare the densities of lumps of material of exactly the same size or volume. Since the volumes are the same, the object with the greatest weight must be the densest, and that with the smallest weight the least dense.

The following table shows the weight of 1 cubic centimetre and 1 cubic foot of a few common solid substances :—

Substance.	Weight of 1 cub. cm.	Weight of 1 cub. ft.
Beeswax . . . .	0·96 gram	60 lbs.
Brass . . . . .	8·2    "	511   "
Butter . . . . .	0·86   "	53   "
Resin . . . . .	1·07   "	67   "
Starch. . . . .	1·53   "	95   "
Sugar . . . . .	1·61   "	100   "
Tallow . . . . .	0·91   "	57   "

**Standard of density.**—But to compare densities it is better to have a standard ; just as we have a standard of length, the yard, with which to compare other lengths, or a standard of size with which to compare other sizes. The density of water at a certain fixed temperature is taken as the standard. (This temperature is called four degrees Centigrade, and written  $4^{\circ}\text{C}$ ., which you will understand after you have studied the thermometer.)

The weight of one cubic centimetre of water at  $4^{\circ}\text{C}$ . is one gram, and its density is taken as the standard of density, and is called 1. Similarly, a substance, of which a cubic centimetre weighs two grams, is said to have a density of 2, and such a substance may be considered to contain twice as much matter as water does, packed into one cubic centimetre.

## DENSITY OF LIQUIDS.

**36. Determination of volume of bottle by the weight of water it holds.**—Procure a small bottle, or medicine phial, with a glass stopper having a groove, made by means of a file, along the part which fits into the bottle (Fig. 23).

A groove of this kind permits liquid to escape when the stopper is put into the neck of a bottle full of liquid. Clean and thoroughly dry the bottle and stopper, and determine their weight in grams. Fill with distilled water. Weigh again. Subtract the weight of the bottle. The remainder gives the weight in grams of the water in the bottle when it is full. Since, as you have learnt, a cubic centimetre of distilled water weighs one gram, you now know the number of cubic centimetres the bottle will hold.



FIG. 23.—Bottle with groove in stopper, for density experiments.

**37. How to find the density of liquids.**—Using the bottle employed in the last experiment, find in succession what weights of methylated spirit, vinegar, salt water, beer, and milk it will hold. Record your results. You know from the last experiment how many cubic centimetres the bottle holds, and you can easily calculate the weight of one cubic centimetre of each of the liquids used, and these numbers give the density of the liquids.

Convince yourself that the same results are obtained by dividing the weight of liquid which fills the bottle by the weight of water the bottle holds.

### **38. Weights of fluid ounces of different liquids.**

—Counterpoise a measuring glass graduated in fluid ounces. Put one fluid ounce of water into the glass. Notice that it weighs one ounce. Pour out the water, and put into the glass one fluid ounce of methylated spirit, turpentine, or mercury. Notice that the one ounce weight no longer balances the fluid ounce of liquid. *Only in the case of water does one fluid ounce weigh one ounce. With other liquids one fluid ounce weighs more or less than one ounce.*

**How the density of a liquid is measured.**—The determination of the density of a liquid is very easy when water is taken as a standard. All it is necessary to do to know the density of any liquid is to ascertain the weight of a cubic centimetre of it. Since the weight of a cubic centimetre of water is one gram,<sup>1</sup> the first weighing shows how many times heavier

<sup>1</sup> At this stage it is assumed that the substance is at 4° C., but this has not been insisted on too much lest the reader should get confused. As a matter of fact, the densities of liquids are generally determined, in actual practice, at 15° C.

or lighter the liquid is than the standard, water ; and this is the density of the liquid.

Density can now be regarded in two ways. It can be considered as

- (1) The number of times a substance is heavier or lighter than water, or, as
- (2) The weight of a cubic centimetre of the substance.

What is true of one cubic centimetre of each of the substances under consideration, namely, the standard water and the substance the density of which is required, will also be true of two cubic centimetres, or any other number of cubic centimetres. So long as the sizes or volumes remain the same, then, we can compare the densities in just the same way, for the number of times the substance is heavier or lighter than water can be determined at once by a simple division sum. But it is not necessary to measure the volume in cubic centimetres. Cubic inches or cubic feet would do just as well. The essential condition for the comparison is that their volumes are the same.

#### **Experimental determination of density of liquids.**

—Since it is only necessary to be quite sure that the volumes of the water and the other substance are exactly the same in order to compare their densities at once, it is a very easy matter to find the density of a liquid. All that has to be done is to find the weight of liquid which will fill a certain bottle, and compare it with the weight of water which fills the same bottle. A convenient way to do this is to use a bottle like the one shown in Fig. 23, having a stopper with a groove cut on it, or a hole bored through it. When such a bottle is filled with a liquid, and the stopper is put in, some of the liquid passes up through the groove to make room for the stopper. A bottle of this kind may be named a *density bottle*.

Suppose the weight of water in a density bottle was found to be 50 grams, and the weight of the same volume of methylated spirit was found to be 40 grams. Then these numbers show the relative densities of the two liquids, and as we take the density of water as the standard or unit, the density of the spirit is equal to 40 divided by 50. It is thus seen that the density of spirit is



represented by the fraction  $\frac{40}{50}$  or  $\frac{4}{5}$ , which written as a decimal fraction is 0·8.

The density of many liquids is greater than that of water. Thus, milk has a density represented by the number 1·03, or  $1\frac{3}{100}$ , so that equal volumes of water and milk would have weights in the proportion of 100 to 103. In all cases we can say

Relative density of substance =  $\frac{\text{weight of substance}}{\text{weight of equal volume of water.}}$

The relative density of a substance is often spoken of as its *specific gravity*.

**Densities of some common liquids.**—Liquids vary very much in density. Among common liquids used in the house some are heavier and some lighter than water. Equal volumes of quicksilver, chloroform, glycerine, blood, milk, and sea-water are heavier than the same volume of pure water, while any quantity of claret, olive oil, turpentine, or petroleum, is lighter than the same volume of water.

## PRINCIPLE OF HYDROMETRY

**39. Some things sink, others float in water**—Put some water into a wide-mouthed vessel such as a fish-globe, and carefully place lumps of different things, *e.g.*, pieces of lead, iron, oak, pine, and cork, one after another, into the water. Observe that (1) some sink and others float, (2) of those which float some sink further into the water than others.

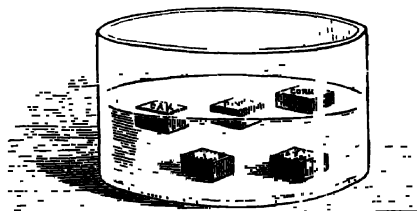


FIG. 24.—When substances are denser than water they sink in water; and if less dense, they float.

**40. Volume of water displaced by solids which float.**

—(a) Obtain a rectangular rod of wood, 1 square cm. in section and about 15 cm. long, with lines around it 1 cm apart. Gouge a small piece of the wood out of one end, and put lead into the hole; flatten the end by filling the hole with wax.

Put some water in a metric graduated jar and notice its level. Find the weight of the rectangular rod, and then place it in the jar with the loaded end downwards. Notice how many cubic centimetres of the rod are immersed, and also how many cubic centimetres of water are displaced (Fig. 25). Since the weight of 1 cub. cm. of water is 1 gram, the number of cubic centimetres of water displaced is also the weight in grams of the water displaced. This weight will be found equal to the weight of the whole rod.

**41. The principle of the hydrometer.**—(a) Fill the graduated glass jar with water up to a certain mark. Notice the level of the water.

Draw lines at equal distances apart across a narrow strip of paper and fix the paper inside a test-tube, as in Fig. 26. Float the test-tube in water in the graduated jar and put shot into it until a certain mark upon the strip of paper inside it is on a level with the surface of the water. Notice the number of cubic centimetres of water displaced when the test-tube is thus immersed.

Then take out the test-tube, dry it, and determine its weight together with the shot it contains. The total weight of the test-tube and contents will be found equal to the weight shown by the number of cubic centimetres of water displaced. Repeat the experiment with the test-tube immersed to a different mark.

(b) Float the test-tube and shot in spirits of wine and milk in succession. Notice that in the former case it sinks deeper than the mark, while in the other not so deep.

**42. Use of hydrometer.**—Place the loaded test tube or a hydrometer (1) in milk, (2) in water, (3) in a mixture of milk and water. Observe the depth to which it sinks in each case (Fig. 27).

**43. Density of brine.**—Make a solution of common salt in water; dissolve salt in the water until a new-laid egg will remain in the position in which it is placed in the solution. Determine the density of the salt solution by the density bottle. The density of the new-laid egg has the same value.

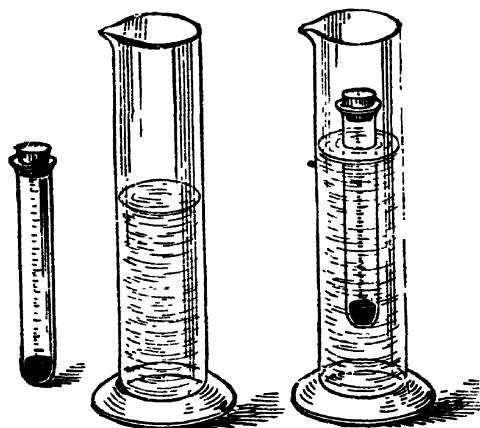


FIG. 26.—Loaded test-tube to illustrate principle of hydrometer.

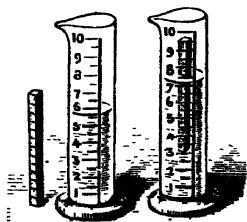


FIG. 25.—Water displaced by floating rod.

**Water displaced by solids which float.**—A solid which sinks in water or any liquid displaces a volume of liquid equal to its own volume. When a solid floats, the case is slightly different. Part of the solid is in water and part

out of the water, and, of course, only the part immersed is pushing the water aside in order to make room for itself. *In the case of*

*a floating object, therefore, the volume of liquid displaced is equal to the volume of the part of the solid below the surface.*

**The depth at which an object floats in water.—**

When any object is floating in water, a certain volume of it is under water and a certain volume is above the surface. The depth at which it floats depends upon its density. A rod of heavy wood sinks deeper in water than a rod of light wood of the same size. The water displaced by the heavy wood has therefore a greater volume, and consequently a greater weight, than that displaced by the light wood. But there is one important fact which applies to both cases : it is that *the weight of the water displaced by the immersed part of a floating object is equal to the whole weight of the object.*

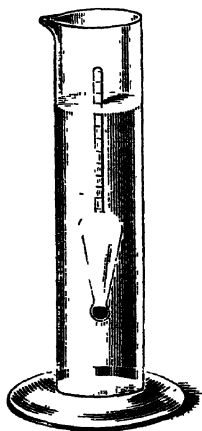


FIG. 27.—Use of hydrometer.

Since an object which floats in water takes a position such that the weight of the water displaced by the immersed part of it is equal to the weight of the object itself, we have a ready way of deciding whether an object will sink further in another liquid or not so far. If the liquid into which it is put is less dense than water, like spirits of wine, it is clear that to make up a given weight we shall want a greater volume of the liquid. Consequently, to displace a weight equal to

the weight of the floating body, the object will have to sink further into the spirit than into the water. If, on the other hand, the object is placed in a liquid such as glycerine, which is denser than water, it will not sink so far, because it will not take so great a volume of this denser liquid to have a weight equal to that of the floating body.

All instruments which depend upon these facts, and are used for measuring the densities of liquids, are called *hydrometers*.

**A familiar instance of hydrometry.—**In making brine for the purpose of pickling meat, a cook often makes use of some of these facts about floating bodies. She knows that a new-laid egg generally has the same density. Though she would not use the same words, she knows that when brine has the same density as a new-laid egg it is of the right strength for

pickling purposes. She goes on adding salt to the water, when she is preparing the brine, until an egg, which she knows to be a new-laid one, will just float in the solution of salt. When this happens, the brine and the egg have the same density, and the amount of salt in the solution is what the cook knows by experience will affect the meat in the way she wants. But it would be better for her to use a hydrometer of some kind, because she cannot be sure that new-laid eggs always have the same density, and it is not very easy to decide when an egg just floats.

**Principle of the lactometer.**—By floating a loaded test-tube in liquids of different densities, as in Expt. 41, several results of great practical importance are obtained. It has been seen that if a mark is made indicating the level at which such a loaded test-tube floats in water, the tube will, if placed in spirits of wine, sink deeper than the mark; while if placed in milk, the test-tube comes to rest with the water mark above the level of the liquid. It can be decided from these facts that the spirits of wine has a density less than that of water, while the milk is denser than water. The construction of a simple instrument, called the *lactometer*, is based upon these conclusions. It is a form of hydrometer employed for measuring the density of milk. When placed in pure milk a lactometer should float with the mark P (Fig. 28) on a level with the surface of the liquid. In a mixture of milk and water the lactometer floats with some other division level with the surface of the liquid. Thus, in milk 10 per cent. below the average density, the 10 above the P mark is level with the surface.

An experienced observer is, therefore, able from the readings of a lactometer to tell whether a sample of milk has a correct density, that is, whether it is heavier or lighter than it should be. At the same time it must be clearly understood that it is not possible to decide at once from the reading of a lactometer whether a sample of milk has been adulterated or not. There are other considerations to be taken into account.

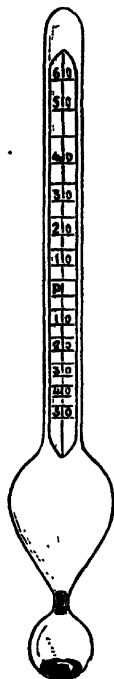


FIG. 28. — A hydrometer for determining if the density of milk is correct. It is called a lactometer.

**The purity of milk.**—Milk may be regarded as a solution in water of certain chemical compounds in which globules of fat are disseminated. The substances in solution include sugar, mineral salts, and a compound of which the student will learn more later, called casein. The presence of more and more of these materials in solution increases the density of the milk. Consequently, if the amount of them present in the milk is increased, the density of the milk is also increased. On the other hand, the density of the fat of milk is less than that of water, with the result that an increase of it in the milk causes the density of the milk to diminish. In estimating the purity of a sample of milk, both these disturbing causes must be taken into account.

If the density of milk is found by a lactometer to be less than it ought to be, it does not follow that water has been added. Some milk is unusually rich in cream, and, containing a high percentage of fat as it does, the milk has a density below the normal value, even though it is especially good. Similarly, a sample of watered milk, to which some sugar or salt has been added, may have the density of ordinary pure milk. In such a case the lactometer would tell us nothing about the malpractices of the dairyman. So that, though in the hands of an analyst a lactometer may give useful information, it would be unwise and unjust to condemn a milkman on the strength of its readings alone.

A good idea as to the richness of milk can be obtained by allowing it to stand in a graduated cylinder until the cream has collected on the top; the proportion of cream is a good index of the value of the milk.

#### SUMMARY.

**Principle of the balance.**—When the pans are at equal distances from the pivot, the beam is horizontal when the weights are equal.

When unequal weights are at different distances from the pivot, there is equilibrium only when

$$\begin{array}{c} \text{weight on} \\ \text{one side} \end{array} \times \begin{array}{c} \text{distance from} \\ \text{pivot} \end{array} = \begin{array}{c} \text{weight on} \\ \text{other side} \end{array} \times \begin{array}{c} \text{distance from} \\ \text{pivot} \end{array}$$

**Standards of weight.**—The *British* standard of weight is the *Imperial standard pound avoirdupois*. The pound is divided into 16 equal parts, each called an *ounce*; the ounce is divided into 16 equal parts, each called a *dram*. A *quarter* is equal to 28 pounds. There are four quarters in a *hundredweight*, and 20 hundredweights in a *ton*.

The *metric* standard of weight is the *kilogram* (1000 grams).

The weight of a cubic centimetre of pure water at 4° C. is *one gram*.

**Density.**—Equal volumes of different materials may have different weights. Equal weights of different substances may have different volumes. Or, substances differ in *density*.

The **relative density** or *specific gravity* of a substance is equal to the weight of any volume of the substance divided by the weight of an equal volume of water.

**Floating bodies.**—When an object floats in a liquid, the *volume of liquid displaced* is equal to the volume of the immersed portion of the object.

The *weight of the liquid* displaced by a floating object is equal to the whole weight of the object.

**The hydrometer.**—The hydrometer is a practical application of the principle of floating bodies. It consists of a suitably loaded tube which is graduated in such a manner that the level at which it floats enables one to tell the density of the liquid in which it floats.

The **lactometer** is a special form of hydrometer used in testing milk. In using it, however, certain other facts, as well as the density, have to be taken into account before pronouncing upon the purity or otherwise of the milk.

### EXERCISES ON CHAPTER III.

1. Write down the names of the weights which are respectively one-tenth, one-hundredth, and one-thousandth of a gram. Also give the names of the weights ten times as heavy as a gram, one hundred times, and one thousand times as heavy.

2. What do you understand by the density of a substance? How would you find the density of olive oil?

3. What is a density-bottle? Explain clearly how you would use it to find the density of sea-water.

4. You are provided with a measuring glass graduated into fluid ounces, a balance, and a set of weights. What observations would you make to find the density of a sample of beer?

5. Arrange the following liquids in order, placing the densest at the top: claret, glycerine, sea-water, and milk.

6. What is the volume of water displaced by an object which floats in it? Describe an experiment you would perform to prove your statement.

7. Will a lead pencil sink further into spirits of wine or into water when floated in these liquids? Give reasons for your answer.

8. A new-laid egg sinks when placed into water, but floats if placed in a strong solution of salt. Why is this?

Does a man float more or less easily in the sea or in a river?

9. What is a hydrometer? How could you find out with a hydrometer which of two vessels contained water and spirits of wine respectively?

10. Draw a lactometer. What precautions have to be taken in testing samples of milk with a lactometer?

## CHAPTER IV

### THERMOMETERS

#### CHANGE OF SIZE

**44. Expansion of a metal bar.**—Place one end of a poker or other bar of metal against a heavy weight, and let the other rest on a needle, as in Fig. 29. Fix a straw upon the needle so that it hangs vertically. Heat the poker with a flame. The straw will be turned to one side on account of the poker as it gets longer turning the needle.

**45. Differential effect.**—Solder a strip or wire of brass, about two feet long, to one of iron of the same length. Straighten the compound

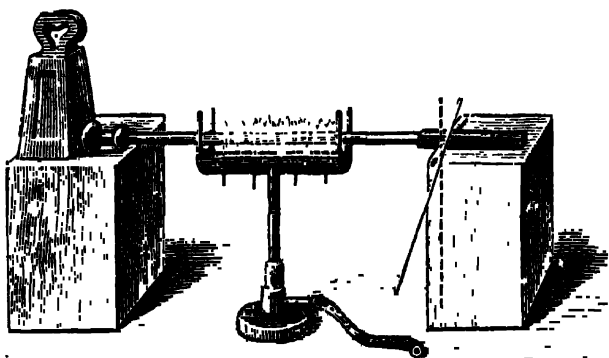


FIG. 29.—Experiment to show the expansion of metal by heat.

strip by hammering; then heat it. Notice that the strip bends, because the brass expands more than the iron. The same effect can be shown, on account of the ebonite expanding more than the wood, by means of a strip of ebonite glued to a strip of wood.

**46. Expansion of liquids.**—Procure a 4-oz. flask and fit it with a cork. Bore a hole through the cork and pass through it a long glass tube which fits tightly. Fill the flask with water coloured with red ink. Push the cork into the neck of the flask and so cause the coloured water to rise up the tube (Fig. 30). See that there is no air between the cork and the water. Now dip the flask in warm water, and notice that the liquid soon gets larger and rises up the tube. Take the flask

out of the warm water, and see that the coloured water gets smaller as it cools, and that it sinks in the tube.

**47. Expansion of air.**—Obtain a flask with a cork and narrow tube. Remove the cork and tube, and, by suction, draw a little red ink into the end of the tube near the cork. Re-insert the cork, and gently warm the flask by clapping it in your hands. Notice that the air in the flask gets larger and pushes the red ink along the tube.

**48. Air thermometer.**—Turn the flask over and place the open end of the tube beneath the surface of some coloured water in a bottle. Warm the flask with the hand, or a flame, so as to expel some of the air, and let the liquid rise in the stem (Fig. 31). (This constitutes an *air thermometer*.)

**49. Air expelled by heat.**—Fit a glass flask with a tight cork through which passes a tube, the upper end of which is bent down and then up at the end. Clamp the flask so that the end of the tube dips under water in a basin. Fill a test-tube with water, and invert it over the end of the tube. Warm the air in the flask, and collect the expelled air in the test-tube.

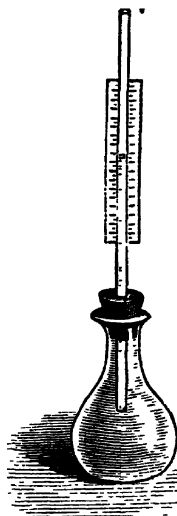


FIG. 30.—To show the expansion of a liquid by heat.

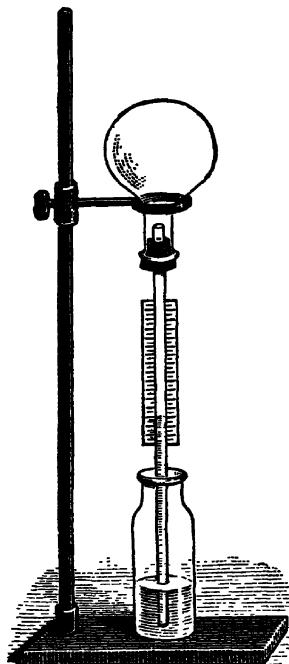


FIG. 31.—An air thermometer.

**Expansion.**—As a rule all bodies, whether solid, liquid, or gaseous, get larger when heated, and smaller when cooled.

The change of size which a body undergoes is spoken of as the amount it expands or contracts; or heat is said to cause expansion in the body. This expansion is regarded in three ways. The expansion of solid bodies may take place in length, in area, and in volume. In the case of liquids and gases we have only expansion in volume.

The expansion which substances undergo when heated, has to be allowed for in many things. Railway lines, for instance,



are usually not placed close together, but a little space is allowed between the separate rails, so that they can expand in summer without meeting. Steam pipes used for heating rooms are also not firmly fixed to the walls at both ends, but are left slightly loose, or are loose-jointed, so that they can expand and contract without doing any damage. For the same reason the ends of iron bridges are not fixed to the supports upon which they rest ; and the bars at the top of a gas stove are left loose, so that they have room to expand when the burners are lighted under them. Iron tyres are put on carriage wheels by first heating the tyre and, while it is hot, slipping it over the wheel. As the tyre cools it contracts and clasps the wheel very tightly.

The common occurrence in domestic life of the cracking of thick glasses when boiling water is poured into them, is to be explained by this expansion of solids by heating. The part of the glass with which the hot water comes in contact is heated and expands, but the heating is confined to one spot, because glass does not allow heat to pass through it readily. It is this local expansion of the glass which results in the cracking of the vessel.

**Measurement of change of temperature.**—Change of temperature means change in the state of hotness or coldness of a body. The change of size which takes place when a thing is heated gives a good way of measuring the change of temperature which it undergoes. For instance, a flask with a narrow neck, or having a cork with a piece of narrow tubing through it fitted into the neck, may contain water up to a certain point marked on the neck or tubing. Suppose this flask of water to be heated by plunging it into warm water, and that the amount by which the top of the water rises in the neck or tube is noted. Then, if the flask is afterwards placed into some other liquid or some more water, and the water is found to rise up the tube to just the same place, it is reasonable to conclude that the second liquid is exactly as hot as the first was. This is measuring its temperature. The flask and tube with the water have become a “temperature measurer,” that is, a *thermometer*.

## TEMPERATURE AND THERMOMETERS

**50. The sense of feeling may be deceived.**—Arrange three basins in a row ; into the first put water as hot as the hand can bear, into the second put lukewarm water, and fill the third with cold water.

Place the right hand into the cold water and the left into the hot, and after half a minute quickly put both into the lukewarm water. The left hand feels cold and the right hand warm while in the same water.

**51. Principle of thermometer.**—Procure an empty thermometer tube, with a bulb at one end and a cup at the top, as in A (Fig. 32), or fix a small funnel upon the tube, as in B. Put some mercury in the cup. Heat the bulb slightly. As the bulb cools, mercury will be forced into the tube to take the place of the air driven out by the expansion of the air in the tube by heat. Repeat the operation until the mercury fills the bulb and part of the stem.

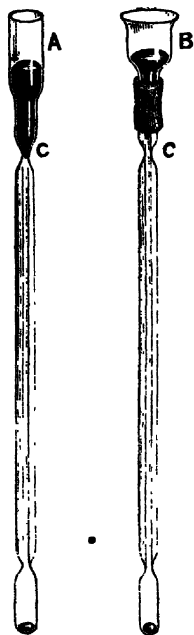


FIG. 32.—Tubes for the construction of thermometers.

**52. Measuring temperature.**—Place in hot water the bulb of the instrument just constructed, and make a mark at the level of the mercury in the tube. Now place the instrument in cold water, and notice that the mercury sinks in the tube. The mercury is thus seen to expand when heated and contract when cooled, and if the glass were marked, the degree of hotness or coldness could be shown by the position of the top of the mercury.

**53. Degrees of temperature.**—Examine a thermometer. Notice that it is similar to the simple instrument already described, but the top is sealed up, and divisions or graduations are marked upon it, so that the height of the mercury in the tube can be easily seen. These divisions are called *degrees* (Fig. 33).

**Feeling of heat and cold.**—

Some people feel cold at the same time that others feel warm. You can therefore easily understand that the sense of feeling cannot be depended upon to tell accurately whether the air or any substance is hot or cold. Some instrument is needed which does not depend upon feeling, and cannot be deceived in the way that our senses can. Such an instrument is called a *thermometer*, and it is used to measure temperature, that is, the degree of hotness or coldness of a body.

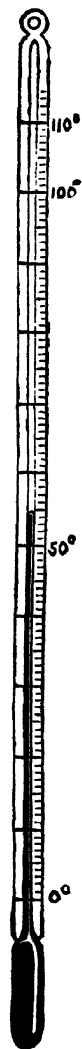


FIG. 33.—A thermometer showing degrees marked on the stem.

**How expansion may indicate temperature.**—You have already learnt that substances usually expand when heated and contract when cooled. A flask filled with water, for instance, and having a stopper through which a glass tube passes, can be used to show the expansion produced by heat and the contraction by cold. But this flask and tube make but a rough temperature measurer. The water does not get larger to the same amount for every equal addition of heat. Neither is it very sensitive, that is to say, it does not show very small increases in the degree of hotness or coldness, or, as we must now learn to say, it does not record very small differences of temperature, and for a thermometer to be any good it must do this. Then, too, as every one knows, when water becomes very cold it is converted into ice, which, being larger than the water from which it is made, cracks the tube. For many reasons, therefore, water is not a good thing to use in a thermometer.

**Choice of things to be used in a thermometer.**—The substance used in a thermometer should expand a great deal for a small increase of temperature. In common thermometers a liquid, either quicksilver or spirits of wine, is employed. Both these liquids expand a fair amount for a given increase of temperature, and, to make this amount of expansion as great as possible, a tube with a very fine bore is used with a much larger bulb containing the liquid.

The liquid in a thermometer must not change into a solid unless cooled very much, nor into a gas unless heated very much. When a thermometer is to be used for measuring great degrees of cold, one containing spirits of wine is employed, since this liquid has to be cooled a great deal before it is solidified. For the measurement of higher temperatures, on the other hand, a mercury thermometer is suitable, because it can be made very hot without being changed into a vapour.

The liquid must be held by a tube, of fine equal bore, which has a large bulb at the end. The bore must be equal all the way along, so that a given amount of expansion in any part of the tube shall mean the same change of temperature; and there must be a large bulb, so that there is a large surface exposed to the substance the temperature of which is to be measured.

There are many other reasons for selecting mercury as the

liquid for an ordinary thermometer. It is a liquid the level of which can be seen easily ; it does not wet the vessel in which it is contained ; it expands a considerable amount for a small increment of temperature ; it is a good conductor of heat, and consequently it very quickly assumes the temperature of the body with which it is placed in contact. Very little heat is required to raise its temperature, and there is therefore very little loss of heat due to warming the thermometer.

**Construction of a thermometer.**—A suitable piece of thermometer tubing having been selected, a bulb must be first blown on one end. The glass is melted at this end and allowed to run together and so close up the bore, and while the glass is still molten, air is blown down the tube from the other end, the tube being kept moving round, so that the bulb is symmetrically placed with reference to it. The bore of the tube is so fine that it is impossible to pour the liquid down it ; some other plan must therefore be adopted. The top of the tube is usually blown into a cup shape as shown at A in Fig. 32, and liquid with which the thermometer is to be filled is put into it. Suppose mercury is being used. Warming the tube makes the air inside it expand, and some is driven out. As the tube cools the mercury is forced into it by the weight of the atmosphere, to fill the place of the expelled air. By repeating this alternate process of warming and cooling, in the circumstances we have described, enough mercury is soon introduced into the tube. The next step is to seal up the tube, leaving no air above the mercury ; to do this the bulb is heated to a temperature slightly higher than that which the thermometer is intended to register, the mercury expands, and when it has reached the drawn-out part, C, of the tube, a blow-pipe flame is directed against it, and the tube is thus closed up. This method of closing a tube and keeping the air out is called *hermetically* sealing it. The thermometer at this stage should be put on one side for some days at least, in order that it may assume its final size, which it does very slowly indeed.

## GRADUATION OF THERMOMETERS.

**54. The temperature of melting ice.**—Take some pieces of clean ice in a beaker or test-tube and plunge a thermometer amongst them. Notice the reading of the thermometer ; it will be either *no degrees* ( $0^{\circ}$ )

or very near it. Warm the beaker or test-tube, and observe that so long as there is any ice unmelted the reading of the thermometer remains the same.

**55. Effect of adding salt to the ice.**—Add salt to the melting ice, and notice that the thermometer indicates a lower degree of temperature.

**56. The temperature of boiling water.**—Boil some distilled water in a flask or test-tube (Fig. 34), with a thermometer a little distance above the surface. Notice that the temperature is  $100^{\circ}\text{C}$ ., or  $212^{\circ}\text{F}$ ., or very near it. Push the thermometer into the water; a higher degree is reached, and the mercury is not so steady as it was in the steam.

**57. Temperature of salt water.**—Add salt to the water. Hold a thermometer in the steam of the boiling water, and notice that the temperature is the same as before, namely,  $100^{\circ}$ . Push the thermometer into the water, and notice that a higher degree of temperature is indicated.

**58. Temperature and changes of state.**—Again place the thermometer in clean ice in a beaker or flask. Gently heat the vessel, and notice the following

changes:—(1) The mercury remains at  $0^{\circ}$  until all the ice is melted. (2) When the ice is melted the mercury rises gradually until it reaches  $100^{\circ}$ . (3) The mercury remains stationary at  $100^{\circ}$  until all the water is boiled away.

**59. Air temperature.**—Notice the temperature of the room indicated by the thermometer.

**60. The clinical thermometer.**—Examine a clinical thermometer. Notice the large space between separate degrees. Observe that just above the bulb the bore of the thermometer is constricted. Hold the bulb in your hand and watch the expansion of the mercury. On exposing the thermome-

ter to the air and allowing it to cool, notice that the mercury thread breaks at the constriction. The mercury can be shaken back into the bulb (Fig. 35).

**The fixed points on a thermometer.**—In the graduation of a thermometer the plan always adopted is to choose

<sup>1</sup> A Centigrade thermometer is supposed to be used. If a Fahrenheit thermometer is used the reading will be  $32^{\circ}$ .

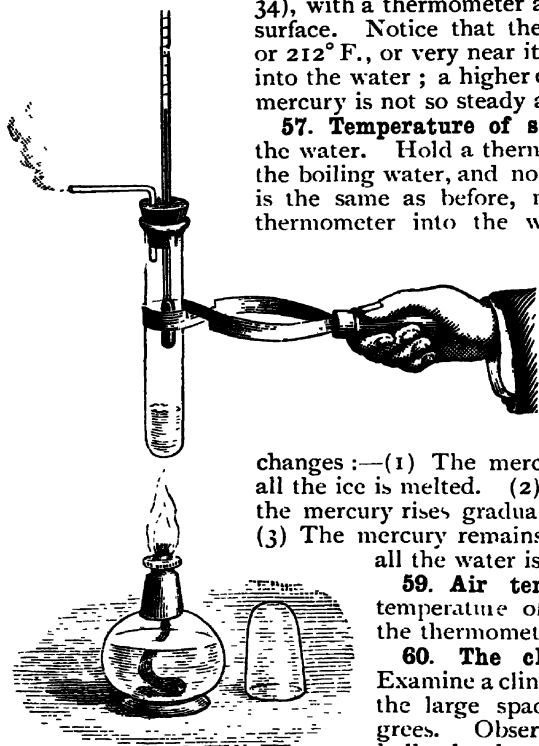


FIG. 34.—How to find the temperature of steam.

"two fixed points" from which to number degrees of temperature. The most convenient lower fixed point is the temperature at which ice melts, or water freezes, for this is always the same if the ice is pure, and remains the same so long as there is any ice left unmelted. Whenever the thermometer is put into melting ice the mercury in it should stand at the same level, or, in other words, melting ice is always at the same temperature and may be thus used to give

one fixed point. The "higher fixed point" chosen is that at which pure water boils at the sea-level. This stipulation must be made, for the boiling point of a liquid is altered when the pressure upon it is changed, being raised if the pressure is greater, and lowered if the pressure is less. When pure water boils, the temperature of the steam is the same as that of the water, and remains so as long as there is any water left. The lower fixed temperature we refer to as the "Freezing Point of Water," the higher as its "Boiling Point."

**Marking the freezing point.**—For this purpose an arrangement like that shown in Fig. 36 is very suitable. The funnel is filled with pounded ice, which before powdering has been carefully washed ; or snow may, if more convenient, be used. The glass dish catches the water which is formed from the melting of the ice or snow. A hole is made in the pounded ice by thrusting in a pencil or glass tube about the size of the thermometer, and into this hole the thermometer is put, and is so supported that the whole



FIG. 35. — A clinical thermometer.

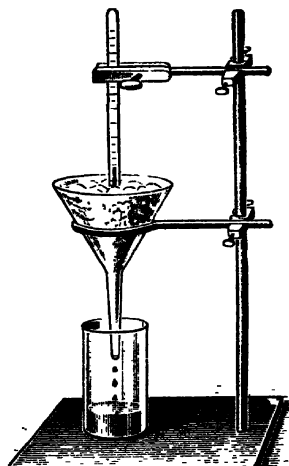


FIG. 36.—Determination of the freezing point of a thermometer

of the mercury is surrounded by the ice or snow. The arrangement is left for about ten or fifteen minutes, until it is quite certain that the tube and mercury are at the same temperature as the melting ice. When this is so, the tube is raised until the mercury is just above the ice, and a fine scratch is made with a three-cornered file on the tube at the level of the mercury.

### Precautions necessary in marking the fixed points.

—As the addition of common salt to ice or snow lowers the temperature of the mixture, great care must be taken that pure, clean ice is used when the lower fixed point of a thermometer is being marked. It must also be noted that the presence of some substances other than common salt has a similar effect on the temperature.

The temperature of a boiling solution of common salt in water is higher than that of the steam which is

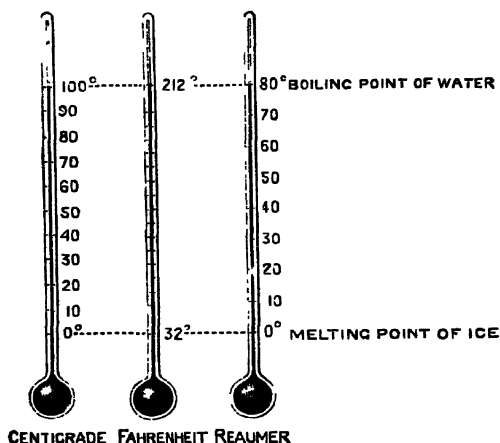


FIG 37.—Thermometric scales.

given off from it. Moreover, since it is the temperature of the steam from boiling water which alone remains constant at a given pressure (the nature of the containing vessel and the presence of substances in solution affecting the temperature of the liquid), it is necessary, when marking the higher fixed point of a thermometer, to see that the instrument is in the steam and not placed in the liquid. It will be understood more fully later that an increase of atmospheric pressure, represented by an increased barometric height, raises the temperature at which water boils; hence the height of the barometer must be recorded when the higher fixed point is being marked.

**Thermometer scales.**—Some value must now be given to the two fixed points which have been obtained as previously described, and of course they may be called anything the

maker of the thermometer likes, but for the sake of comparing one man's observations and experiments with those of other people, it is most convenient to graduate all thermometers upon an established plan. The thermometers used in this country are divided up in two ways—(1) the Centigrade scale, (2) the Fahrenheit scale. A third scale—the Réaumur scale—is extensively used in Germany (Fig. 37).

**The Centigrade scale.**—Here the freezing point is called *zero* or *no degrees Centigrade*, written  $0^{\circ}$  C. The boiling point is called *one hundred degrees Centigrade*, and is written  $100^{\circ}$  C. The space between these two limits is divided into 100 parts, and each division called a *degree Centigrade*.

**The Fahrenheit scale.**—On thermometers marked in this way the freezing point is called *thirty-two degrees Fahrenheit*, written  $32^{\circ}$  F., and the boiling point *two hundred and twelve degrees Fahrenheit*, written  $212^{\circ}$  F. The space between the two limits is divided into 180 parts, and each division is called a *degree Fahrenheit*. The physicist Fahrenheit, after whom the thermometer is named, took as the beginning of his scale the point at which the mercury stood when his thermometer was placed in a mixture of ice-water and either sal-ammoniac or sea salt. He called this point zero, or  $0^{\circ}$ , and the point at which the mercury stood when placed in pure ice and water,  $32^{\circ}$ . Fig. 38 shows equivalent degrees of temperature on the two scales.

**Domestic uses of thermometers.**—It is of the first importance that one or more thermometers should be in every household, ready to hand, and frequently used.

For the measurement of the temperature of the body, what is termed a *clinical thermometer* is best (Fig. 35). As the temperature of the living human body is never many degrees above or below a temperature of  $98^{\circ}$  F., a clinical thermometer is only graduated from about  $95^{\circ}$  F. to  $110^{\circ}$  F. If the bulb

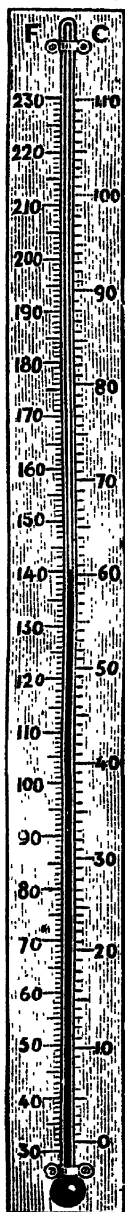


FIG. 38. — Fahrenheit and Centigrade degrees.



of such a thermometer is put into the mouth, or under the arm-pit, of a person in health, and left there for two or three minutes, it will be found, on taking it out, to indicate a temperature from  $97^{\circ}\cdot 8$  F. to  $98^{\circ}\cdot 6$  F. The thread of mercury in the stem of the thermometer remains in one position, though the air is cooling the mercury while the thermometer is being read. This is because of the constriction at the top of the bulb, which causes the thread of mercury in the stem to be left behind while the mercury in the bulb contracts. To "set" the thermometer for a fresh observation, it is only necessary to jerk it slightly, when the thread of mercury will again join up to the liquid in the bulb.

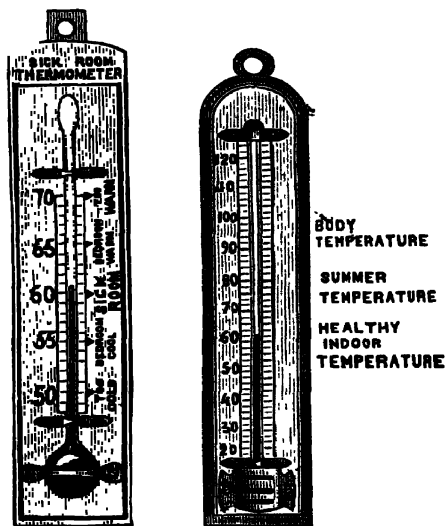


FIG. 39.—House thermometers.

The body may feel hot or cold at different times, but its actual temperature varies only very slightly, whether a person in good health is at the hottest or coldest parts of the earth, eating or fasting, at rest or taking violent exercise. In a fever, however, the temperature of the patient may rise to  $105^{\circ}\cdot 8$  F., or even higher.

The temperature of a living-room should be as

nearly as possible  $60^{\circ}$  F.; and to indicate it a thermometer should be kept in a part of the room free from draughts, and away from the fire or other source of heat.

It is always advisable to test the temperature of the water of a bath, and for this purpose a bath thermometer, which is surrounded with a wood casing to enable it to float, is convenient. Baths may be roughly classified as follows: cold bath,  $60^{\circ}$  to  $70^{\circ}$  F.; tepid bath,  $85^{\circ}$  to  $95^{\circ}$  F.; warm bath,  $96^{\circ}$  to  $104^{\circ}$  F.; hot bath,  $104^{\circ}$  to  $114^{\circ}$  F. A warm bath for a child should have a temperature of  $96^{\circ}$  to  $98^{\circ}$  F. The water of a cold bath should

never have a temperature less than 59° F., that is, about forty degrees lower than the temperature of the healthy human body.



FIG. 40.—Bath thermometer.

Few kitchens are provided with a thermometer. Yet many of the operations of cooking can only be regulated with that degree of nicety which ensures success when the help of a thermometer is called in. For instance, the clear, gum-like contents of an egg, known as albumin, become converted into a white solid, as in the white of a hard-boiled egg, at a temperature of 60° C. (140° F.). Butter, if pure, melts at about 95° F., and margarine at about 86° F. If the cook adopts the unscientific plan of guessing, or "trusting to luck," we too often have an uncooked or hard-boiled egg, burning fat, and other unpleasant troubles, which are the more vexing since they could so easily be avoided.

Just as every cook, who knows her business, makes constant appeals to the pair of scales which she always has handy, so she frequently finds it necessary to call in the assistance of a thermometer.

**Graphic plan of showing temperatures.**—The seven illustrations of a thermometer in Fig. 41 show the temperature indicated by a certain thermometer at noon

on the first seven days of a certain month. A dotted line has been drawn from the point at which the top of the mercury

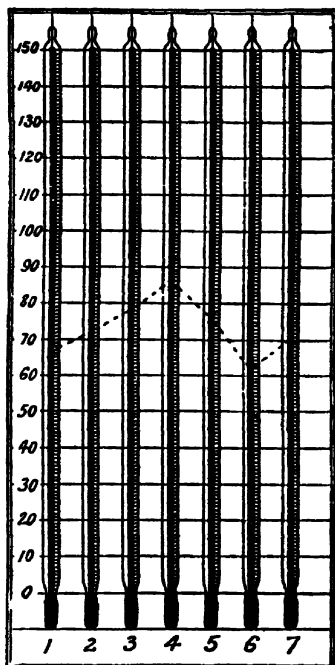


FIG. 41.—Temperatures on seven successive days at noon.

stood on each day to that which it occupied the next day. The line is, in fact, a graphic representation of the changes of temperature from day to day. But it is not necessary to draw the thermometer, nor even the thread of mercury; it is enough

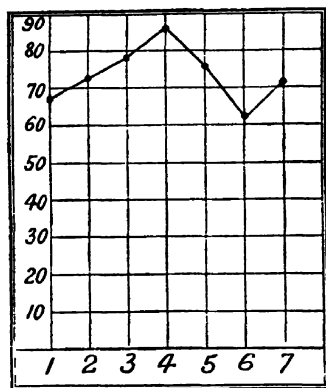


FIG. 42.—Graphic representation of temperatures at noon on seven successive days.

to put a dot at the point where the top of the mercury stands day by day, as shown in Fig. 42, and then the dots can afterwards be connected by a line, which will clearly indicate the rise and fall of temperature.

### SUMMARY.

**Expansion.**—As a rule all bodies, whether solid, liquid, or gaseous, become larger when heated, and smaller when cooled.

Expansion and contraction by heat must be taken into account in (a) laying railway lines, (b) fixing steam or hot-water pipes, (c) building bridges.

Contraction is utilised in fixing iron tyres to wheels.

**Choice of materials to be used in a thermometer.**—(1) The substance used should expand a great deal for a small increase of temperature.

(2) If a liquid is used it should not change into a solid unless cooled very much, nor into a gas unless heated very much.

(3) The liquid should be in a fine tube with a comparatively large bulb at the end.

**Reasons why mercury is used for thermometers.**—In addition to the above reasons, (a) its level can be easily seen, (b) it does not wet the vessel in which it is contained; (c) it expands a considerable amount for a small increment of temperature; (d) it is a good conductor of heat; (e) very little heat is required to raise its temperature.

**The fixed points on a thermometer.**—(1) The temperature at which ice melts or water freezes; (2) the temperature of the steam issuing from boiling water when the barometer stands at 30 inches.

**Thermometric scales.**—The distance between the freezing and boiling points on the stem of a thermometer may be divided as follows:

	Fahrenheit scale.	Centigrade scale.
Boiling point.....	212°	100°
Freezing point....	32°	0°

**Other important temperatures.**—The temperature of the healthy human body is about 98° F. [97°·8 to 98°·6].

In fever the temperature of the body may rise to  $105^{\circ}\cdot 8$  F. or even higher.

The temperature of a living-room should be as nearly as possible  $60^{\circ}$  F.

*Bath Temperatures.*

Cold bath  $60^{\circ}$ – $70^{\circ}$  F.      Warm bath  $96^{\circ}$ – $104^{\circ}$  F.

Tepid bath  $85^{\circ}$ – $95^{\circ}$  F.      Hot bath  $104^{\circ}$ – $114^{\circ}$  F.

Warm bath for child  $96^{\circ}$ – $98^{\circ}$  F.

EXERCISES ON CHAPTER IV.

1. What is a thermometer, and what information concerning heat does it supply?

2. How would you test whether the two fixed points on a mercurial thermometer were accurately marked?

What would be the effect on the values obtained for the fixed points of the presence of common salt as an impurity in the materials used in determining these points? (1898.)

3. A flask containing pure water is heated by a single burner and one thermometer is placed with its bulb below the surface of the water, and another thermometer with its bulb just above the surface. When the water boils the readings of the two thermometers are taken. Will the readings be the same?

What will be the effect on the reading of each thermometer (1) of placing a second burner under the flask, and (2) of dropping some common salt into the flask? (1898.)

4. Describe carefully the manner in which the freezing and boiling points on a thermometer are determined. (1897.)

5. Take a glass tube open at one end and having a bulb at the other. Hold the tube so that the open end dips into water. Heat the bulb gently with a spirit lamp for a minute or two, and then take the lamp away. What will be observed? How can you account for the facts observed? (P.T., 1898.)

6. Mention the conditions that should be satisfied by the bulb and tube of a mercury thermometer. Give the reason for each condition that you state. (P.T., 1898.)

7. Describe carefully how to construct a common thermometer.

8. Write an essay pointing out the uses of a thermometer in the household.

9. Draw and describe a clinical thermometer. What temperature will it show after its bulb has been placed in the mouth of a healthy man?

## CHAPTER V

### CAPACITY FOR HEAT. CHANGE OF STATE

#### HEAT AND TEMPERATURE

**61. Distinction between temperature and heat.**—Place a can containing water over a burner. Place in the can a small test-tube containing water. After the can has been heated for a little time, observe the temperature of the water in the test-tube and in that surrounding it ; it will be the same. Take away the burner, and lift the test-tube out of the can. You now have a small quantity of water and a larger quantity both at the same temperature ; but there is more heat in the large amount than in the small amount. Prove this by pouring the hot water from the test-tube, and that from the can, into equal quantities of cold water from the tap in separate vessels. The large amount of hot water will thus be found to have a greater heating effect than the small amount ; hence it must have possessed more heat than the small amount.

**62. Result of mixing equal hot and cold quantities of the same substance.**—Put a certain weight of warm water in a glass, and the same weight of cold water in another glass. Observe the temperature or each by means of a thermometer. Pour the cold water into the hot. It will be found, on stirring them together with the thermometer (taking care not to break it), that the temperature of the mixture is about midway between the two original temperatures.

**Difference between heat and temperature.**—Temperature is not heat, but an effect produced by heat ; it is only a state of a body, for the body may be cold one minute and hot the next. A hot body is one at a high temperature, a cold body one at a low temperature. If a hot body and a cold body are brought into contact, there is an exchange of heat until they are both of the same degree of hotness or coldness, that is, at the same temperature. Hence, temperature may be defined as *a condition or state of a body which is changed by the gain or loss of heat.*

**Changes of temperature when hot and cold liquids are mixed.**—Suppose that a certain weight of hot water is put

into one vessel and an equal weight of cold water into another. There will then be equal weights of water at different temperatures. If the two liquids are mixed together, the temperature of the hot water will fall, and the temperature of the cold will rise. The loss of heat of one will be equal to the gain by the other, so that the temperature of the mixture will be midway between the two original temperatures. Thus, if the weights of water are equal, and the temperatures at first are  $60^{\circ}$  C. and  $20^{\circ}$  C., then the temperature of the mixture will be  $40^{\circ}$  C. The temperature of the hot water will fall  $20^{\circ}$  C., and the temperature of the cold water will rise  $20^{\circ}$  C.

The actual temperature of the mixture will be slightly less than the calculated temperature, because some heat will be lost while the liquids are being mixed.

**Quantity of heat in water at different temperatures.**—Quantity of heat may be measured by heating effect, so that we can say that the quantity of heat in a vessel of water depends upon the *weight* of the water and its *temperature*. For any temperature, say  $60^{\circ}$  C., the amount of heat in 100 grams of water is twice as great as in 50 grams of water. Similarly, the amount of heat in 100 grams of water at  $40^{\circ}$  C. may be regarded as twice that in 100 grams at  $20^{\circ}$  C.; assuming for the sake of simplicity that water at  $0^{\circ}$  C. contains no heat. The fall of temperature multiplied by the weight of hot water is equal to the rise of temperature multiplied by the weight of cold water.

**Unit quantity of heat.**—As in all other cases of measurement, a unit or standard quantity is required with which to compare quantities of heat. The unit quantity of heat generally adopted is the *amount of heat necessary to raise the temperature of one gram of water through one degree Centigrade*. This unit is called a *calorie* or *therm*. The amount of heat required to raise the temperature of 2 grams of water through  $1^{\circ}$  C. is thus 2 units or 2 calories. Similarly, if 1 gram of water at  $0^{\circ}$  C. is heated in a test-tube over a burner until its temperature is  $1^{\circ}$  C., it will have received from the burner 1 unit of heat, or 1 calorie. If the tube contain 10 grams of water at  $0^{\circ}$  C., and its temperature is raised to  $12^{\circ}$  C., it will have received 10 times 12 ( $=120$ ) units of heat, the number of units being equal to weight (in grams)  $\times$  increase of temperature (in degrees Centigrade).

## HEAT CAPACITY

**63. The same quantity of heat may produce different changes of temperature.**—Weigh out equal quantities of water and turpentine at the same temperature in two vessels of the same size. Pour equal quantities of hot water at the same temperature into the cold water and the turpentine. Observe the rise of temperature produced in each case. Though the equal amounts of hot water contain the same quantity of heat, the rise of temperature of the turpentine will be found to be more than the rise of temperature of the cold water.

**64. Comparative heating effects of lead and water.**—Place equal weights of lead and water in two test-tubes standing in the same vessel of water, and heat them over a laboratory burner until the water boils; the temperature of both the lead and the water will then be about  $100^{\circ}$  C. Provide two glasses containing equal weights of cold water at the temperature of the room. Carefully (in order not to break the glass) put the hot lead into one of these and the hot water into the other. Stir both mixtures and note the temperature in each case. The water into which the heated lead is plunged is not at so high a temperature as that into which the hot water is poured.

Equal amounts of water at the same temperature are thus shown to be heated to different degrees by equal amounts of lead and water at the same high temperature.

**65. Capacity for heat.**—Place some iron nails in a glass, and the same weight of cold water in another glass. Let the two glasses stand for a while so as to assume the temperature of the room. Boil water in a kettle or other vessel, and pour equal quantities into the two glasses. Observe the temperature of the mixture in the two glasses. The iron nails will be found to be hotter than the water in the other glass, because iron takes less heat to raise its temperature than is required by an equal weight of water at the same temperature.

**Comparison of heat quantities.**—The quantity of heat in any amount of water depends upon the weight of the water and its temperature. It might be supposed, therefore, that as any weight of water at a certain temperature contains a certain quantity of heat, the same weight of another substance at the same temperature contains the same quantity of heat. This is not the case. 100 grams of water at a temperature of  $50^{\circ}$  C. contain 5000 units of heat more than at  $0^{\circ}$  C., but 100 grams of turpentine, mercury, lead, iron, or any other substance at the same temperature as the water, namely  $50^{\circ}$  C., do *not* contain this additional number of units of heat. The quantity of heat in a substance thus depends not only upon the weight and the temperature, but also upon the substance itself.

If some hot water is poured into a cup containing cold water,

the cold water rises to a certain degree of warmth ; but if the same weight of hot water is poured into a cup containing the same weight of cold turpentine and at the same temperature as the cold water, the mixture of water and turpentine becomes warmer than the mixture of hot and cold water, because turpentine has a smaller *capacity for heat* than water.

The heat capacity of a substance may be considered as analogous to the capacity of a vessel to hold water, and temperature to change of level, as illustrated in Fig. 43.

**Capacity of water for heat.**—Of all known substances, water has the greatest capacity for heat. Consequently a larger amount of heat is required to raise the temperature of a given weight of water through any number of degrees than is needed by an equal weight of any other substance.

Thus, suppose a pound of water be put into one flask and a pound of mercury into another, and that these flasks are then heated for

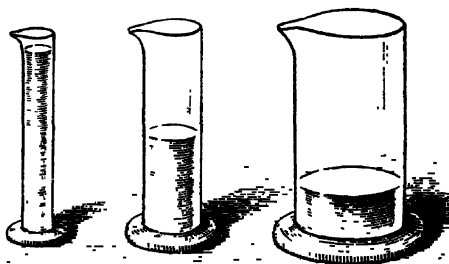


FIG. 43.—The same quantity of water causes different rises of level in vessels of different capacities.

five minutes by two laboratory burners, which, so far as appearances go, give out the same quantity of heat. The temperature of the two liquids at the commencement of the experiment is, say,  $15^{\circ}$  C. If at the end of the experiment the temperature of the water is  $20^{\circ}$  C., that of the mercury would probably be about  $180^{\circ}$  C., that is, much higher than the boiling point of water.

The capacity for heat of a substance is sometimes called its *specific heat*.

**Results in nature of high capacity of water for heat.**—The results in nature of this great capacity for heat which water possesses are very important.

Though water takes a large amount of heat to warm it, and is consequently heated by the sun's rays only slowly, yet when it cools it parts with its heat just as slowly. The effect of this on the climate of islands is very marked. The winter temperature



is never very low, and the climate never very severe, because the water surrounding the country acts as a great storehouse of heat, slowly giving up heat to it. Similarly, the summer temperature is never unbearably hot, because the surrounding water takes so long to warm, and, being always cooler than the land, continually takes heat from the latter.

## LIQUEFACTION

**66. Melting point of butter.**—Place a little butter in a test-tube, and stand a thermometer in it. Place the test-tube in a beaker of water, being gently heated on a sand-bath. Notice the temperature at which the butter melts. Take out the test-tube when the butter has all melted and let it cool. Notice the temperature at which the butter solidifies.

**67. Melting point of ice.**—Put some small pieces or shavings of clean ice into a beaker, and thrust a thermometer into them. Record the temperature indicated. Pour in a little water, stir the mixture, and again record the temperature. Place the beaker on a sand-bath and warm it gently. Notice the reading of the thermometer *so long as there is any ice unmelted*. In all these cases the reading of the thermometer is practically the same, indicating that the temperature of melting ice is constant.

**68. Heat required to melt ice.**—Let a few lumps of ice stand in a beaker until some of them have melted. Notice that the temperature is  $0^{\circ}$  C. Counterpoise two empty beakers of the same size in the pans of a balance, and put a small lump of the ice into one, and the same weight of water from the melted ice in the other. You have thus equal weights of ice and water at  $0^{\circ}$  C. Pour equal weights of hot water into the two beakers. When the ice is melted, observe the temperature of the water in each beaker. The temperature of the water in the beaker in which the ice was placed will be found much lower than that of the water in the other beaker, owing to the ice using up a large quantity of the heat in melting into water.

**69. Comparative cooling effects of ice and water.**—Take equal weights of hot water in two large beakers of the same size. Place a piece of ice in one of the beakers, and observe the temperature of the water when it has melted. Pour ice-cold water into the other beaker until the same temperature is reached. Find the weights of ice and ice-cold water which have been added. It will be found that a small weight of ice has as much cooling effect as a large weight of ice-cold water.

**70. Heat given up during solidification of a liquid.**—Dissolve as much crystallised sodium sulphate as you can in boiling water contained in a flask. Transfer the *saturated* solution, as it is called, to another flask which has been thoroughly cleaned. Gently put a thermometer into the flask, and close its neck with a plug of cotton wool. Place the flask and its contents on one side to cool. When the solution is of the same temperature as the room, it is what is known as a *supersaturated*

solution. Now drop a small crystal of sodium sulphate into the liquid. Notice that the solution suddenly solidifies ; observe the rise in temperature as indicated by the thermometer.

**Change of state.**—Substances exist in three states, namely, solid, liquid, and gaseous. By the action of heat a substance may be changed from one state to another. Wax, for instance, is usually a solid, but by heating it becomes a liquid. Butter can in the same way easily have its state altered from solid to liquid. Lead and zinc are also melted when heated, but they require a hotter flame than wax or butter.

A good example of the changes of state produced by heat is obtained by heating a piece of ice until it becomes water, and then heating the water until it passes off into steam or water vapour. Here the same form of matter is by heat made to assume three states ; in other words, ice, water, and steam are the same form of matter in the solid, liquid, and gaseous state respectively.

**Temperature of melting.**—When a solid is heated, the first effect is usually an increase of size. But if the heating is continued long enough, when the solid reaches a certain temperature, which differs for different solids, melting begins. The solid changes into a liquid. The temperature at which the melting takes place is called the *melting point*. Thus, when a lump of lead is heated, its temperature rises, it gets larger, and as the heating is continued it is converted into a silvery-looking liquid. Wax, ice, and iron are other examples of solids which melt. But ice, wax, lead, and iron differ very widely in the temperatures at which they begin to melt, as the following table shows :

Ice	melts at . . . . .	0° C.
Bees-wax	„ . . . . .	65° C.
Lead	„ . . . . .	330° C.

So long as any of the solid remains unmelted, the temperature does not rise above the melting point.

**Latent heat.**—It is certain that when a mixture of ice and water is heated over a laboratory burner, heat is being continually given to the mixture. Yet the temperature as recorded by the thermometer gets no higher. What becomes of this heat, as it has no effect upon the temperature of the mixture ? You know that the ice is gradually melted, and if the heating is continued

long enough it is all changed into water. So soon as this has happened, every further addition of heat raises the temperature of the water. These considerations lead to the conclusion that the heat previously given to the mixture is all used up in bringing about the change of ice into water. Further, it is found that not only in the case of ice, but when any solid is turned into a liquid, there is no increase in temperature, even while heat is being added to it, until the whole of the solid has been changed to a liquid. This amount of heat which is necessary to change a solid into a liquid is spoken of as *latent heat*.

Just as in the change of a solid into a liquid at the same temperature, a definite quantity of heat becomes latent, or used up, so when the converse change occurs a definite amount of heat is liberated. This is well shown by preparing a supersaturated solution of sodium sulphate, and then causing it to solidify suddenly by dropping a crystal of the same substance into it. When the change from the liquid to the solid state is thus brought about, the heat, which in the liquefaction became latent, is liberated again, and a decided rise in temperature, which immediately affects the thermometer, is noticed.

**Latent heat of water.**—The number of units of heat which are required to change the state of a gram of ice, converting it from the solid to the liquid condition without raising its temperature, is called the latent heat of water, or the *latent heat of fusion of ice*. To melt 1 gram of ice requires 80 heat-units. That is to say, as much heat as would raise the temperature of a gram of water through  $80^{\circ}\text{C}$ ., or would raise that of 80 grams of water through  $1^{\circ}\text{C}$ ., is used up in changing a gram of ice into a gram of water at the same temperature.

**Natural consequences of latent heat of water.**—Just as it is necessary before a pound of ice can be changed into a pound of water to supply an amount of heat which would raise the temperature of a pound of water through  $80^{\circ}\text{C}$ ., so before a pound of water can be changed into a pound of ice, we must take from it precisely the same amount of heat. This is why it takes so many cold nights to cover a pond with ice; for not until every pound of water at the surface has lost this amount of heat can it change into ice. For the same reason, it takes a very long time to melt completely the snow in the roads and the ice on the ponds, even after a thaw has set in.

## VAPORISATION

**71. Cooling produced by evaporation.**—Sprinkle a few drops of spirits of wine, sal volatile, and ether, on your hand in succession. Notice the liquid soon disappears, and its presence in the air can be detected by its smell. The rate at which the liquid evaporates is increased by waving the hand about. The hand feels cold.

**72. Water frozen by evaporation.**—Pour a few drops of water upon a dry piece of thin wood, and stand in the water a thin beaker containing a little ether. Blow vigorously down a tube having one end in the ether, or use a pair of bellows (Fig. 44).

The ether rapidly evaporates, and *in doing so takes heat from the water* between the beaker and piece of wood. The beaker and wood become frozen together.

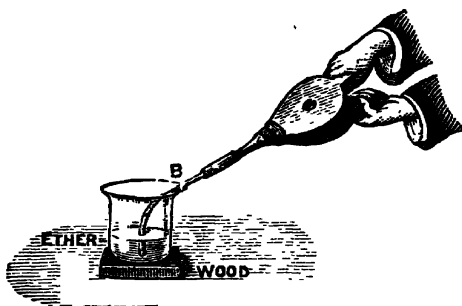


FIG. 44.—How to freeze water by the evaporation of ether.

**73. Heat liberated when steam is condensed.**—Put equal quantities of cold water into two beakers of the same size. Observe the temperature. Boil water in a flask with a delivery tube (Fig. 45), and pass the steam into the cold water in one of the beakers. When the temperature of the water has been raised about ten degrees, take away the delivery tube, pull the cork out of the flask, and pour enough boiling water into the other beaker to raise the temperature by the same number of degrees. Find by means of a balance the weights of steam and water added to the cold water in the two beakers. It will be found that a small weight

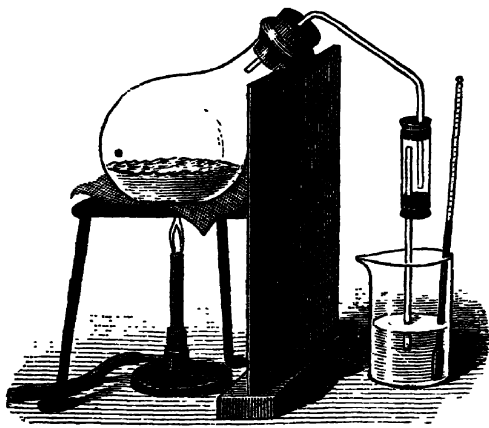


FIG. 45.—Flask fitted with delivery tube and trap to catch condensed steam. For latent heat experiment.

of steam in condensing into water produces the same heating effect as a much larger weight of boiling water.

**Heat disappears during vaporisation.**—When a liquid is changed into vapour, a certain amount of heat is used up. It does not matter whether the liquid evaporates quietly or boils; every gram of it requires a certain amount of heat before it becomes converted into vapour. In boiling, this heat is supplied by the flame or fire, and in evaporation it is taken from the objects in contact with the liquid. The faster the evaporation the more heat is absorbed in this way. When a liquid evaporates very rapidly, the cooling produced is very noticeable. For instance, if a few drops of either spirits of wine or ether are sprinkled upon the hand, the liquid soon disappears and the hand feels cold. The heat necessary for the evaporation of these liquids is taken from the hand (or from any other things with which they are in contact), consequently the hand becomes cooler and cooler as the vapour is formed. So much heat may be absorbed in this way that water can be frozen by the evaporation of ether in a vessel in contact with it.

In tropical countries, where the land gets *very hot* during the day, evaporation takes place so rapidly after sunset that the water sometimes becomes so much cooled as to freeze, by the extraction of the heat required to bring about the change from liquid to vapour.

**Latent heat of steam.**—When once water has started to boil, its temperature gets no higher than the boiling point. So long as there is any water left, no matter how much you heat it, its temperature remains the same. All the heat is absorbed, or used up, in bringing about the change from the liquid state to that of vapour. It requires a great many more heat-units to convert one gram of water at a temperature of  $100^{\circ}$  C. into steam at the same temperature, than it does to change a gram of ice at  $0^{\circ}$  C. into a gram of water at  $0^{\circ}$  C. It has been seen that to bring about the latter change requires an expenditure of 80 heat-units, whereas to convert a gram of water at  $100^{\circ}$  C. into a gram of steam without changing its temperature requires no fewer than 536 heat-units. Thus the *latent heat of steam*, or, as it is sometimes called, the *latent heat of vaporisation of water*, is 536. It must be remembered that a liquid is never changed into a vapour without some absorption of heat. This is true whether the change takes place quietly in evaporation, or rapidly as in boiling.

Just as a large quantity of heat is required to convert water into steam, so a large quantity is given up when steam becomes water.. It is for this reason that a scald from the steam of boiling water is worse than a scald from the boiling water itself.

### SUMMARY.

**Temperature** may be defined as a condition, or state, of a body which is changed by the gain or loss of heat.

**The unit quantity of heat** is the amount of heat necessary to raise the temperature of one gram of water through one degree Centigrade. This unit is called a *therm*, or a *calorie*.

To ascertain the number of heat-units taken up, or given out, by any weight of water we may say :

Number of heat-units = Weight of water in grams  $\times$  Number of degrees Centigrade through which its temperature rises, or falls.

**The capacity for heat** of a substance is analogous to the capacity of a vessel for fluids. Water has a greater capacity for heat than any other substance. Its high capacity for heat has a profound influence on the climate of islands.

In determining the number of heat-units in a substance, its capacity for heat must be taken into account as well as its weight and its temperature. In fact :

Number of heat-units = Weight of substance  $\times$  its change of temperature  $\times$  its capacity for heat.

**Melting point.**—The temperature at which a solid changes into a liquid is called its *melting point*.

**Latent heat.**—The number of heat-units necessary to change a solid into a liquid at the same temperature, or the amount required to change a liquid into a gas at the same temperature, is known as *latent heat*.

**Latent heat of water.**—The number of heat-units required to convert one gram of ice at  $0^{\circ}$  C. into water at the same temperature is known as the *latent heat of water*. Its numerical value is 80.

**Latent heat of steam.**—The number of heat-units required to change one gram of water at  $100^{\circ}$  C. into steam at the same temperature is known as the *latent heat of steam*. Its numerical value is 536.

### EXERCISES ON CHAPTER V.

1. If a pound of water at  $100^{\circ}$  C. is mixed with a pound of water at  $0^{\circ}$  C., the temperature of the mixture is  $50^{\circ}$  C. How would the result have differed if a pound of oil at  $100^{\circ}$  C. had been substituted for the hot water? Explain the difference. (1897.)
2. What is the difference between temperature and quantity of heat?

What experiments would you perform to show that equal weights of water and iron tacks when cooled through the same range of temperature give out different quantities of heat?

3. What is the capacity for heat of a body?

Which has the greater capacity for heat, 5 c.c. of mercury, or 2 c.c. of water? [Specific gravity of mercury, 13.6; specific heat, .033.] (Junior Oxford Local, 1900.)

4. What will be the temperature of the mixture produced (supposing no heat to be lost) by mixing—

(a) 1 lb. of water at  $10^{\circ}\text{C}$  with 3 lbs. of water at  $70^{\circ}\text{C}$ ?

(b) 1 lb. of quicksilver at  $15^{\circ}\text{C}$ . with 1 lb. of mercury at  $95^{\circ}\text{C}$ ?

5. I place on a piece of wood a vessel containing ether, having previously poured some water on the wood. State what occurs when I blow over the surface of the ether with a pair of bellows, and explain the several results. (1900.)

6. Define what is understood by the melting point of a solid. How would you determine this temperature in the case of butter?

7. What is meant by the latent heat of fusion? How do you account for the fact that a long time is necessary to effect a complete thaw after a hard frost?

8. What unit is employed in measuring quantities of heat?

9. Why is a scald from the steam of boiling water more severe than one from boiling water itself?

## CHAPTER VI

# TRANSFERENCE OF HEAT

### CONDUCTION

**74. Relative conductivities of metals.**—Obtain wires of copper, iron, brass, German silver, and of any other metals available. Let the diameters be the same as nearly as possible, and the lengths about 15 to 20 cms. Place the wires upon a clay tile, as shown in Fig. 46. Support the tile in a horizontal position, and heat the wires with a flame where they meet. After a few minutes, slowly move the head of a match along each wire in succession, commencing at the ends away from the flame, and notice the points at which the matches will light.

The distances of these points from the flame indicate the relative conducting powers of the metals of which the wires are composed.

**75. Lowering of temperature by conduction.**—Make a short coil of stout copper wire  $\frac{1}{4}$ -inch internal diameter. Pass it over the wick of a lighted candle without touching the wick. The candle is extinguished owing to the cooling effect of the wire, which conducts away the heat.

**76. Action of wire gauze.**—Turn on, but do not light, a gas jet. Hold over it a piece of wire gauze, and light the gas above the gauze. Notice that the flame does not strike through. Why? Vary the experiment by lowering a piece of cold wire gauze upon an ordinary Bunsen flame. What happens?

**77. Water is a bad conductor.**—Fill a test-tube three-quarters full with cold water. Hold the test-tube near the bottom, and warm the top of

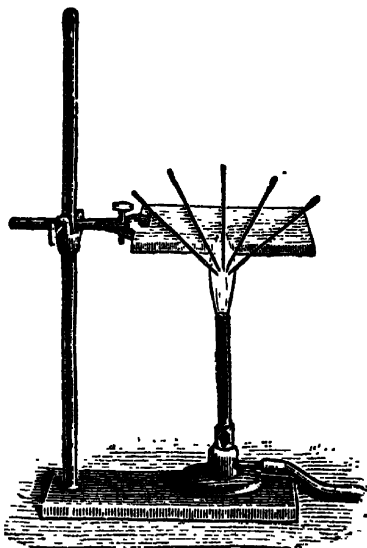


FIG. 46.—Experiment to show that heat travels along some metals more easily than along others.



the water in a Bunsen flame, as shown in Fig. 47. The water at the top can be heated until it boils vigorously, but if the test-tube is taken away from the flame, and the temperature of the water at the bottom tested with a thermometer, it will be found much below the boiling point.

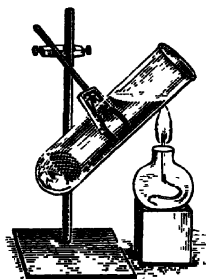


FIG. 47.—Experiment to show that water is a bad conductor of heat.

**Conduction of heat.**—By touching a succession of things in a room, say the marble mantel-piece, the fender, the back of a chair, the hearth-rug, a succession of sensations is obtained; the first two feel cold, the chair-back not quite so cold, while the rug feels quite warm, and yet they are one

and all under the same conditions, and there is no reason why they should not be at the same temperature. The explanation of these different sensations is really very simple. In all those cases where the hand *receives more heat than it gives up* the sensa-

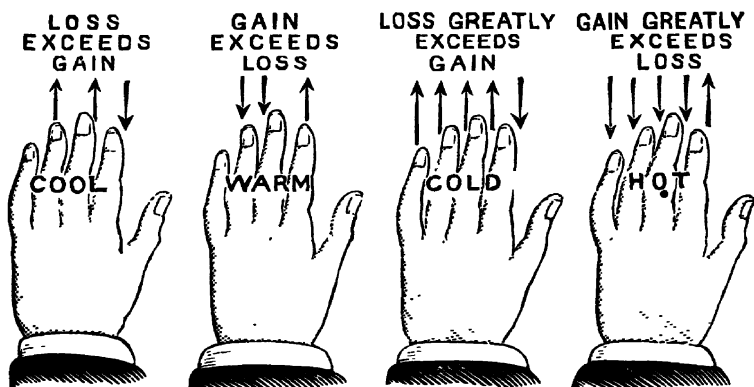


FIG. 48.—Relation of gain of heat to loss decides whether the hand feels hot or cold.

tion of warmth is experienced, while in those where the hand *loses more heat than it gains* the body feels cold or cool (Fig. 48). This explains why the fender feels colder than the hearth-rug. The fender takes more heat from the hand than the hearth-rug, and it does so because it is a better *conductor of heat*.

It is worth while to consider this expression a little. Suppose one end of a poker is put into a fire, while the other is held in the hand. Soon the poker begins to feel warm, and as time goes on it gets warmer and warmer, until at last it can be held no longer. Heat has passed from the fire along the poker, or has been *conducted* from the fire by the poker.

*The process by which heat passes from one particle of a body to the next is called conduction, and the body along which it passes is known as a conductor.*

**Good and bad conductors.**—Those substances which easily transmit heat in this way are called *good conductors*, while those which offer a considerable amount of resistance to its passage are called *bad conductors*.

Metals are, as a rule, good conductors of heat, but some metals conduct heat better than others. Most liquids are bad conductors of heat, though mercury, being a metal, is an exception. If liquids were heated only by conduction, water would boil throughout just as quickly when the source of heat was placed in contact with the top layer of liquid as it does when the heating takes place from below.

Gases are even worse conductors of heat than liquids. The very small power of conducting heat possessed by air is made use of in many familiar household devices.

**Everyday uses of bad conductors.**—To keep ice in the warm days of summer the custom is to wrap it up in flannel and put it into a refrigerator. The flannel, because of its loose texture, encloses a quantity of air which, being a bad conductor of heat, prevents the passage of heat from the warm outside air to the cold ice inside. The refrigerator, too, depends for its construction on much the same facts. It is often a double-walled box with an air space between the walls, which forms a non-conducting layer retarding the passage of heat from outside; or, the space may be filled with *asbestos*, or some other bad conductor.

Ice which has to be conveyed by rail or boat is often packed in sawdust. The reason for this is because the air between the particles of wood making up the sawdust does not conduct the heat from the outside to melt the ice. For the same reason, to lift hot plates it is usual to hold them in a folded cloth, when the layer of air enclosed by the cloth prevents the hand being burnt.

The handles of metal teapots are often made of ivory, bone, or similar non-conducting material, to enable the teapot to be lifted without inconvenience, even when it is full of hot tea. A tea cosy, inclosing as it does a layer of air, keeps the tea hot, because the escape of heat from the teapot takes place only slowly.

One of the chief reasons for wearing clothes is because they are bad conductors of heat. In winter they keep in the heat of the body, so that it cannot escape to the cold outside air; and in summer they keep the heat of the sun from the cooler body. Such devices as ladies' muffs and woollen comforters all owe their efficacy to the air inclosed by them, which, because of its being a non-conductor of heat, prevents the heat of the hand and neck from escaping, and so causing them to become cold. For a similar reason blankets and eider-down quilts are used on the bed in cold weather.

The wooden handle of a soldering "iron," and the cork mat on which one stands in a bath-room, are both used to avoid an escape of heat. Evidently wood and cork are bad conductors of heat.

## CONVECTION

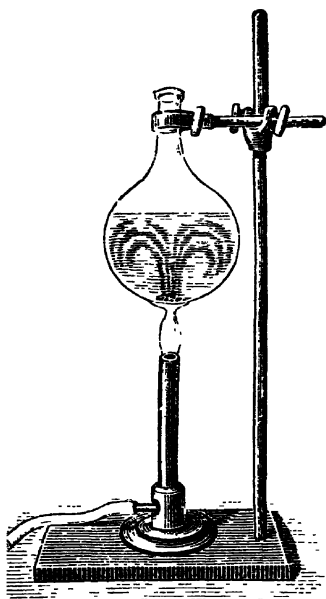


FIG. 49.—Convection currents in a liquid.

**78. Convection in a liquid.**—Heat over a small flame a round-bottomed flask full of water, as in Fig. 49. Throw into the water some solid colouring matter, like

cochineal, aniline dye, litmus, etc. Notice how the hot and coloured water ascends.

**79. Circulation of water.**—Fit up the apparatus shown in Fig. 50. A is a 6-oz. wide-mouthed corked bottle, with the bottom knocked out (a small gas jar will do, or an ordinary lamp glass may be used). A well-fitting cork with two holes is inserted, through which the bent glass tubes B, B' pass, as shown. They are united at the bottom by a short piece of india-rubber tubing, C. Pour water into A until it just covers the open ends of the tubes. Now pour in about a tea-spoonful of ink. Apply a small flame at B. Notice what happens.

### 80. Convection currents in gases.—

Place a short piece of candle in a saucer, light it, put a lamp glass over it, and pour sufficient water into the saucer to cover the bottom of the lamp glass (Fig. 51). Watch how the light of the candle is affected. Next cut a strip of card less than half the height of the lamp glass, and nearly as wide as the internal diameter of the top. Insert the card into the lamp glass so as to divide the upper part into halves. Now light the candle again, and see whether it will burn with the divided chimney over it. Test the direction of the currents of air at the top of the chimney by holding a smoking taper or lighted match over it.

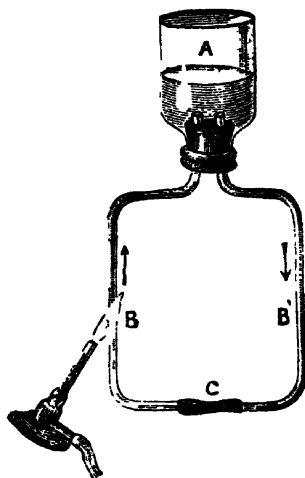


FIG. 50.—Circulation of water caused by heat.

**Process by which liquids are heated.**—The process by which water and other liquids are heated may be easily studied by heating water into which bran or oak dust,

or some solid colouring matter, has been thrown in a round-bottomed flask over a small flame, as in Fig. 49. The water nearest the flame gets heated, and consequently expands and gets lighter. It therefore rises, and causes a warm ascending

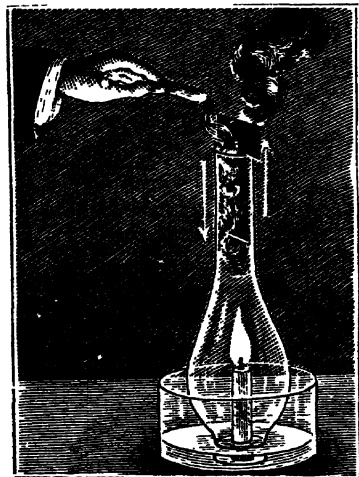


FIG. 51.—Convection currents in air.

current of coloured water. But something must take the place of this water which rises, and the cold water at the top, being heavier than the warm water, sinks to the bottom and occupies the place of the water which has risen. This water in its turn gets heated and rises, and more cold water from the surface sinks. This gives rise to upward currents of heated water and downward currents of cool water, until by and by the whole of the water is heated. These currents are known as *convection currents*, and the process of heating in this manner is called

*convection.* Eventually the whole of the water gets so hot that the bubbles of vapour which are formed near the source of heat are not condensed again in their upward passage through the liquid, and coming to the surface they escape as steam.

Gases are similarly heated by the process of convection, which may be thus defined :—*Convection is the process by which fluids (liquids and gases) become heated by the actual movement of the particles of the fluid due to difference of density.*

The upward movement of warm air and downward movement of cooler air can be shown by placing a short piece of lighted candle in a saucer, putting a lamp glass over it, and pouring sufficient water into the saucer to cover the bottom of the lamp glass (Fig. 51). In this case the light of the candle is affected and eventually goes out. But if a strip of card is cut less than half the height of the lamp glass, and nearly as wide as the internal diameter of the top, and is inserted into the lamp glass, so as to divide the upper part into halves, and then the candle is again lit, it will be seen to continue to burn with the divided chimney over it. The direction of the currents of air at the top of the chimney can be shown by holding a smoking taper or smouldering brown paper over the chimney.

## RADIATION

**81. Transmission of heat by radiation.**—Place two similar thermometers, one about a foot away from the flame of a laboratory burner, and the other eighteen inches away. Arrange the thermometers and the flame in one straight line. Notice that the bulb of the thermometer nearer to the flame is hotter than the one more remote. How does the heat of the flame travel to the thermometer?

Place one of the thermometers eighteen inches above the flame, and compare the readings of the thermometer in this and its former position. In the latter case the air is heated by convection.

**82. Distinction between radiation and conduction.**—(a) Fix a metal wire with one end in the flame of a laboratory burner. The wire becomes hot. After a few minutes slowly move the head of a match along the wire until it ignites. Notice the distance at which this occurs. Bring another match near the side of a flame and notice how close it must be to the flame before it ignites. Evidently air does not conduct heat so well as a metal.

(b) Light a burner, and hold one hand by the side of the flame as close as you can to it; then suddenly turn off the gas with the other hand. The hand ceases to be heated so soon as the flame is extinguished, thus showing that the air around the flame is not appreciably warmed by

radiation. A metal wire, however, held in the flame remains hot even after the flame has been extinguished.

**83. Effect of surface upon radiation.**—Obtain two small bright tin cans or canisters, and fit into each a cork having a hole through which a thermometer will pass. Cover the outside of one of the vessels with lamp-black by holding it over a candle or luminous gas flame, or over burning camphor. Put the same quantity of hot water at the same temperature in each, and then cork up the vessels, each cork having a thermometer through it so that the bulb is well immersed in the water. Observe the temperature of each vessel of water, and if the temperature of one is higher than that of the other, cool the vessel until the temperatures are equal. Then put the vessels in a cool place where there are no draughts, and after 20 to 30 minutes again read the temperatures.

The blackened vessel will be found to have lost or radiated more heat than the bright one.

**84. Effect of surface upon absorption.**—Similarly fill equally a blackened and a bright vessel with cold water of the same temperature, and hang them for 20 to 30 minutes at the same distance above an iron plate, supported on a tripod stand and heated by a laboratory burner. At the end of this time observe their temperatures.

The blackened vessel will be found at a higher temperature than the bright one, indicating that a lamp-black surface absorbs heat better than a bright metallic surface.

**Radiant heat.**—Heat can travel from one place to another by radiation, which is distinct from conduction or convection. The respects in which radiation differs from the other ways in which heat moves from one place to another are : (1) it travels in straight lines, and (2) it does not warm the medium through which it travels. Although you may not have thought of the fact that radiation travels in straight lines, you have made use of it when you have screened your face from the heat of the sun or of a fire. When you wish to protect yourself from the glare of the summer sun you seek a shady place, because then some object, it may be a tree or a house, is in the straight line between you and the sun.

Rays of heat are, in fact, similar to rays of light, and both are transmitted in the same way. It is easy to show that heat-rays do not appreciably warm the air through which they are passing. For instance, if the hand is held close to one side of the flame of a laboratory burner, the hand feels warm, but almost immediately the flame is extinguished the sense of hotness ceases, whereas, if the air by the side of the flame were actually heated by the rays, it would feel hot even when the flame was extinguished. In this respect the air behaves differently from a

metal rod, one end of which is heated in a flame, for the rod remains hot after the flame has been turned out, because it is heated by conduction.

**Radiation and absorption.**—When two tin cans, one of which is very bright on the outside, while the other is coated with lamp-black, have equal quantities of hot water placed in them and are left to cool in a place free from draughts, it is found that the water in the blackened can gets cooler in a given time than that in the bright can.

Similarly, if equal quantities of cold water are placed in a bright and a lamp-black-covered can, respectively, and the cans are hung above a hot metal plate, it is found that after a given interval of time the water in the blackened vessel is hotter than that in the bright vessel.

Substances of different kinds do not absorb and radiate heat equally. Bright metal surfaces, for instance, are bad absorbers of radiant heat, but they are excellent reflectors of radiation. Lamp-black is a good absorber of radiant heat, and it, in common with other good absorbers, radiates heat with great readiness. It is a general rule that good absorbers are good radiators.

**Applications of principles.**—These facts about radiators and absorbers are constantly being made use of in everyday life. It is customary to wear lighter coloured garments in summer than in winter ; the reason is that the darker clothes are better absorbers of heat than lighter ones, though, since they are also better radiators, it is very questionable if there is much efficacy in the custom. The reason usually alleged for wearing light-coloured clothes in summer is that they are good reflectors of the radiant heat from the sun. But in this case, too, it must be borne in mind that these light fabrics are also bad radiators, and consequently the heat of the body is less readily got rid of ; but, as the outside air is much warmer than the body this does not detract very much from their comfort. It would probably be more satisfactory to wear light-coloured clothes both in summer and winter ; the temperature of the body is not, however, the only consideration. In a climate like ours the questions of cleanliness and economy have also to be taken into account. Where expense is no object, the best results are obtained from light-coloured garments for both seasons.

The bad absorbing and good reflecting powers of white things

are made use of, too, in the white sunshades which are carried by ladies in summer, and in the custom of white-washing the roofs of buildings to make the rooms inside cooler during the hot days of July and August.

Bright metal tea-pots are preferred because of the bad radiating powers of their surfaces. The tea in such a tea-pot will keep hot longer than that in a black earthenware one if they are both left on a table for the same length of time. On the contrary, if the tea has become cool, and it is required to warm it up, by placing the tea-pots in front of the fire, the tea in the black tea-pot will become heated more quickly because the black surface absorbs heat more readily than the bright metal. For a similar reason stoves and fire-grates are made black so that they may the more readily radiate their heat.

#### SUMMARY.

**Heat is transmitted** in three ways: (1) by conduction; (2) by convection; (3) by radiation.

**Conduction** is the process by which heat passes from one particle of a body to the next.

Solids are usually better conductors than liquids; liquids conduct heat better than gases.

**Everyday uses of bad conductors.**—The plan of wrapping ice in flannel, the construction of refrigerators, the tea-cosy, eider-down quilts, are all applications of the bad conducting power for heat of air.

Cork and wood are also bad conductors of heat.

**Convection** is the process by which liquids and gases become heated by the actual movement of their particles due to difference of density.

Heating buildings by hot water and methods of ventilation are applications of heating by convection.

**Radiation** differs from conduction and convection in two ways: (1) it travels in straight lines: (2) it does not warm the medium through which it travels.

**Radiation and absorption.**—Different substances do not absorb and radiate heat equally. Good absorbers are good radiators. Lamp-black is a good absorber and a good radiator. Bright metal surfaces do not absorb heat much, but reflect it readily.

The custom of wearing light-coloured garments in summer is due to the fact that they do not absorb so much heat as darker ones. Roofs are white-washed for a similar reason.

#### EXERCISES ON CHAPTER VI.

I. What is meant by convection?

Illustrate your answer by sketches, taking the case of a vessel filled with water and heated from below, and explain why it is that convection is set up. (1899.)



2. Why is a vessel of water heated more quickly if heat is applied at the bottom than if it is heated at the top?

Draw a diagram to illustrate the movements of a liquid heated from below. (1897.)

3. Point out the difference between the conduction and convection of heat.

Describe an experiment showing that water is a bad conductor of heat. (1897.)

4. Water sometimes spurts from the spout of a kettle standing upon a fire. How do you account for this, and how would you prevent it without taking the kettle off the fire. (P.T., 1897.)

5. On a cold morning a gardener grasps the iron part of his spade with one hand and the wooden part with the other. Explain why one hand feels colder than the other. (P.T., 1898.)

6. If a spoon made of solid silver and one made of brass and only silver-plated are placed in bowls in some boiling water, the handle of the silver spoon becomes much hotter than that of the plated one. Why is this?

Describe an experiment by which you would show that your explanation is correct. (1899.)

7. How is the reading of a thermometer altered by wrapping a wet rag round the bulb? What will happen if the rag is wetted with (1) ether, (2) oil instead of water? How do you explain the various results? (Queen's Sch., 1899.)

8. Two test-tubes A and B are filled with water. A small piece of ice is allowed to swim in A, and a similar piece of ice is sunk by a weight to the bottom of B. Heat is applied to the closed end of A and to the open end of B. In which test-tube may we expect the ice first to melt? and in which may we expect the water first to boil? Give reasons for your answer. (Queen's Sch., 1899.)

9. What is the use of a tea cosy? To what properties of heat does it owe its efficacy? Mention any other familiar applications of the same facts.

10. Describe an experiment, the results of which will enable you to compare the conducting powers of copper, German silver, and iron for heat.

11. Explain how the air of a room is heated by an ordinary fire burning in it. Compare this state of things with what goes on in nature, resulting in the formation of winds.

12. How does the heat of the sun reach the earth? Compare this process with that by which the air in a room is heated by the ordinary open fire.

13. Which do you consider more suitable for winter wear, dark or light coloured clothes? Give reasons for your answer.

14. In which kind of tea-pot will the tea keep hot longer, a bright silver one or a black earthenware one? Why?

15. Why do ladies sometimes screen their faces by holding a fan between themselves and the fire? What does this teach about the way in which radiant heat passes through a space?

## CHAPTER VII

### SOLUTION, EVAPORATION, DISTILLATION

#### SOLUTION OF SOLIDS

**85. Soluble and insoluble solids.**—(a) *Sugar*.—Place a piece of sugar in water; note that it soon disappears, and gives a sweet taste to the whole of the water, so that in some way the sugar must have spread throughout the water.

(b) *Sand*.—Add sand to water, and stir it up with the water. Let the water stand for a short time, and notice that the sand sinks to the bottom.

(c) *Camphor*.—Stir up camphor with water. Notice that the camphor

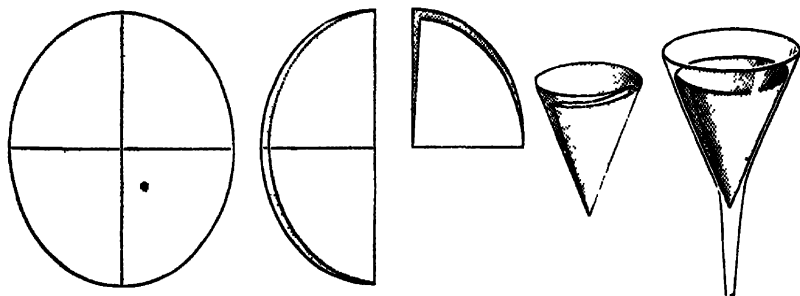


FIG. 52.—How to use a round piece of blotting-paper for filtering.

does not disappear; it is insoluble in water. Shake up a small lump of camphor with some spirits of wine, in a small bottle. It gradually disappears, just as sugar does in water.

(d) *Sulphur*.—Shake up in a bottle flowers of sulphur with carbon bisulphide, and notice that it disappears. *Be careful to keep the stopper in the bottle of carbon bisulphide, and do not bring the bottle near a light.*

**86. Removal of suspended substances by filtration.**—Fold a circular piece of clean white blotting-paper, or a filter-paper, in the manner shown in Fig. 52. Insert the folded paper into a glass funnel, and place the funnel in the neck of a flask. Make some water muddy by stirring mud

into a tumbler of water, or by putting powdered charcoal into it. The mud or charcoal remains suspended in the water for a long time.

Pour the muddy water carefully on to the filter-paper in the funnel, in the manner shown in Fig. 53, and observe that the water which drops through is quite clear. The mud is left on the paper.



FIG. 53.—How to pour a liquid into a filter paper in a funnel.

**87. Non-removal of dissolved substances by filtration.**—Similarly, filter a solution of sugar or salt, and observe that the solution is unaltered by passing through the paper.

**What “soluble” and “solution” mean.**—When sugar is stirred up in a cup of tea, it disappears. By tasting the tea, it is possible to find out if much sugar has been put into the tea or not. Similarly, sugar gives a sweet taste to water and milk; if salt is added instead, an entirely different taste is given to the liquid.

Whenever a substance

disappears in a liquid in this way, and yet can be recognised by suitable means everywhere in the liquid, it is said to dissolve or form a solution, while the substances which disappear are said to be soluble. Washing-soda, borax, and saltpetre are all easily soluble in water.

**Substances insoluble in water.**—Many things, on the contrary, will not dissolve in water, and these are spoken of as insoluble. Sand, gravel, coal, camphor, are all instances of substances insoluble in water.

But though camphor will not dissolve in water, yet it disappears when shaken up in spirits of wine. And as camphor is

soluble in spirits of wine, a solution of camphor in spirits of wine can be made. Shellac is another substance which will dissolve in spirits of wine and not in water ; such a solution, in fact, makes a varnish used for covering some kinds of furniture. Sulphur, again, though it will not dissolve in water, and only very little in spirits of wine, disappears very quickly if placed in the nauseous liquid called carbon bisulphide. Chalk will not dissolve in water, but it will dissolve in a weak acid, such, for instance, as vinegar.

**Substances held in suspension in water.**—Substances which are insoluble in water will, if finely powdered and stirred up with water, often take a long time to settle ; that is, the fine particles remain *suspended* in the water for a long time. The lighter the particles are, the longer the time it takes for them to settle and for the water to become clear. If on a rainy day a glassful of muddy water is taken from the gutter, and then placed on one side, the particles can be watched settling to the bottom. Those substances which, like the mud, are spread throughout the water without being dissolved in it are said to be *held in suspension*. The rate at which these suspended particles settle to the bottom to form a *sediment* depends upon their density. Light particles take a long time, heavy particles only a short time to sink.

**Suspended substances can be removed by filtering.**—It is easy to separate suspended impurities from water. The process by which this is done is called *filtration*. Many substances are used through which to filter water containing particles in suspension. Chemists most commonly use paper which has not been glazed ; such paper is porous ; the holes through it are large enough to let water pass, but not large enough to let the suspended substances go through. In consequence, these particles are left on the paper in the funnel, and the water which trickles through is quite clear. It must be remembered, however, that it is impossible to get rid of substances *in solution* by filtering the liquid. Dissolved material passes through the holes in the paper with the liquid in which it is held in solution.

Other substances besides unglazed paper are used sometimes in filtering. Thus, often the water supply of a town is filtered through beds of sand. Household filters are made with pieces

of charcoal for the water to trickle through, and in some others a particular kind of porous iron or porcelain is employed. Every filter requires to be cleaned frequently, or it gets clogged with impurities from the water which has filtered through it.

## SOLUTION OF LIQUIDS

**88. Mixture of alcohol and water.**—Pour some water into a bottle and then some alcohol,<sup>1</sup> and shake them up together. Observe that the alcohol disappears in the water or dissolves in it.

**89. Liquids which do not mix.**—Shake up together some olive oil and water, and allow the mixture to stand for a short time. Notice that the liquids separate into two layers, the lighter being on the top. Which is the lighter?



**90. Gas dissolved in water.**—Examine a bottle of soda-water. Notice that it appears clear and bright, and seems to have nothing dissolved in it. Uncork, or otherwise open it. Bubbles of gas escape (Fig. 54). A lighted taper held to the mouth of the bottle has its flame put out by the gas which is given off.

**Solution of one liquid in another.**—The commonest cases of solution are when solid substances are dissolved in liquids. In addition to this, many liquids will dissolve in other liquids. Alcohol and glycerine are examples of liquids which dissolve in water. When a liquid dissolves in water, as alcohol and glycerine do, we say, in ordinary language, that the two liquids *mix*. This is only another way of saying that they form a solution. When the amount of alcohol is greater than that of the water, it is very common to say that the alcohol has been *diluted* with water.

**Liquids insoluble in one another.**—If, however, oil, water, and mercury are shaken up together, and then left to stand for a time, they will be found to separate from one another and lie in different layers—the mercury at the bottom, oil at the top, and water between the two (Figs. 55 and 56). Here, then, we have examples of liquids which do not dissolve in one another or do not mix. Oil and water will not mix, quicksilver and water will not mix, and so on.

<sup>1</sup> Ordinary methylated spirit will not do, as it forms a milkiness with water. If pure spirits of wine cannot be obtained, whisky or brandy will do.

**Some gases dissolve in liquids.**—When a bottle of soda-water is opened, bubbles of gas rise out of it. The gas has evidently been dissolved in the liquid. This is only one of many instances of gases which will dissolve in liquids. There is a large amount of this gas, which is known to chemists as carbon dioxide or carbonic acid gas, dissolved in soda-water. Many other gases dissolve in water; the liquid sold by chemists as liquid ammonia is a solution of ammonia gas in water.

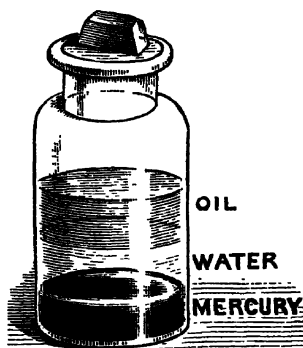


FIG. 55.—Liquids which do not mix.

**Importance of air dissolved in water.**—Rain in falling through the air dissolves some of it in its passage to the earth. The air thus dissolved in the water serves an important purpose. Both animals and plants must have air to breathe. As is well known, some animals and plants live in water, and these depend upon the air which is dissolved in the water. When water is boiled, the dissolved air which it contains is driven out of it by the heat.

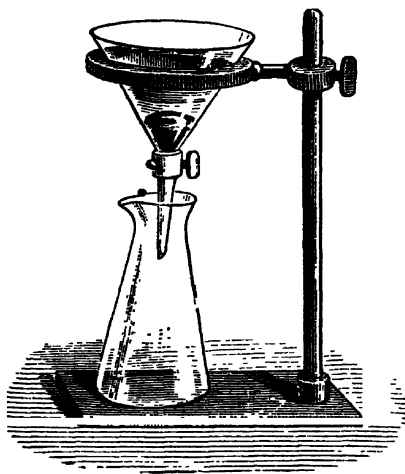


FIG. 56.—A funnel with a tap, for drawing off the liquids in it one after the other.

## OTHER SOLUTIONS

**91. Saturated solutions.**—Procure a supply of alum (or of nitre) and powder it. Put some of the powdered solid into a flask and add water. Shake them up together for some time, and if all the powder dissolves add more and shake again. Continue this addition of the powder and the shaking until some powder remains undissolved, however much it is shaken. A *cold* saturated solution, that is, a solution containing as much of the solid as it will hold is procured.

Now warm the cold saturated solution. The powder which before remained at the bottom of the flask dissolves. Continue to add more

alum, or nitre, and notice that a great deal must be added before you obtain a *hot* saturated solution.

**92. Formation of crystals.**—Place the hot saturated solution on one side to cool. As cooling proceeds, some of the alum or nitre separates out in clear, well-formed crystals, because as the solution cools it cannot dissolve so much alum as before.

If time permit, make saturated solutions of salt, washing-soda, borax, and lime.

**93. Relative solubilities.**—Find the weight of a small evaporating basin. Into it, by means of a pipette, put 10 cubic centimetres of the saturated solution of alum or nitre previously made. Place the basin upon a sand-bath and gently heat the liquid, which will gradually disappear, leaving the solid behind in the basin. When the residue is quite dry, allow the basin to cool, and weigh it again. The increase in weight tells the amount of alum or nitre dissolved in 10 cubic centimetres of the saturated solution.

Find out, in the same way, the amount of solid in 10 cubic centimetres of each saturated solution prepared.

Represent your results graphically as explained on p. 81.

**94. Solution in acids.**—Procure a lump of marble, and break it into small fragments, or powder some chalk (*not* blackboard chalk). Satisfy yourself it will dissolve neither in cold nor hot water. Pour a few drops of vinegar or weak hydrochloric acid upon some of the fragments in a test-tube. Observe the fizzing which begins; place a lighted match at the mouth of the test-tube, the flame is put out. After a time the fragments of chalk or marble will have disappeared.

**Saturated solutions.**—When any given amount of water has dissolved as much of a solid as it can be made to, without warming or assisting it in any other way, it is said to be saturated. But though cold water, for instance, may be saturated with any particular solid, such as sugar, it can, if warmed, be made to dissolve more sugar. Though there are some exceptions, it may be regarded as the general rule that water and other liquids will dissolve more of a solid when they are warm than when they are cold. In some cases the amount of solid which will dissolve goes on increasing as the temperature of the water is increased. In general, therefore, the cooler the water the less of a solid will it dissolve. Now suppose warm water is given as much sugar, salt, alum, or any substance of this kind as it will hold, and is then cooled, what happens? It has to give up some of the substance, for it cannot hold so much as when it was warm. When tea has been made very sweet, some of the sugar is left on the bottom of the cup as the tea cools. This is because, though the tea was able to dissolve a certain

amount of sugar when hot, it could not hold so much when cold, and therefore a little of it was deposited upon the bottom of the cup.

**Graphic Representation of Solubilities.**—The relative solubilities of different substances may be shown conveniently in a graphic manner. It has been found by experiment, for example, that alum, nitre, soda, and common salt, are dissolved in boiling water and in cold water to the extent shown in the following table :—

100 parts of cold water dissolve	100 parts of boiling water dissolve
14.5 parts . Alum	300 parts
30     "     . Nitre	200     "
50     "     . Soda	110     " (about)
36     "     . Salt	37     "

Now, if on a piece of squared paper we arrange that one side of a square shall represent 10 parts of the substance, we can mark off lines of lengths proportional to the numbers shown in the above table, when we shall obtain a diagram like that shown in Fig. 57, which shows at a glance the relative solubilities of the different substances in cold and boiling water respectively.

**A change sometimes accompanies solution.**—In all the instances of solution already examined, the solid has quietly disappeared in the liquid to form a clear solution : but though some solids are quite insoluble in water and spirits of wine,

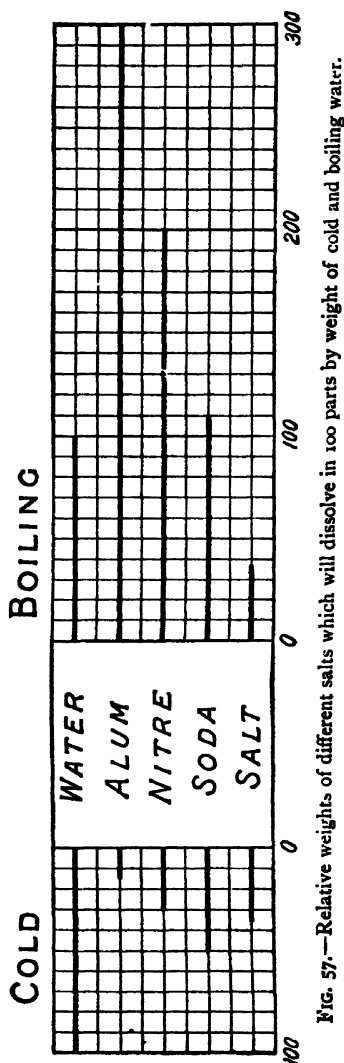


Fig. 57.—Relative weights of different salts which will dissolve in 100 parts by weight of cold and boiling water.



they can be made to dissolve when treated with liquids called acids. Marble and chalk are instances of this. Fragments of either of these substances disappear neither in water nor in spirits of wine, but if they are acted upon with weak hydrochloric acid, or a little vinegar, both of which liquids are acids, a brisk fizzing takes place, and the pieces of marble soon pass into solution. The fizzing means that a gas is being given off; the truth of this statement is easily proved by putting a lighted taper to the mouth of the vessel in which the solution is taking place when the flame is extinguished. As will be fully seen in later lessons, this experiment forms a good example of what is called a chemical change, to distinguish it from those other cases which have been studied, called physical changes, where solution is brought about without an obvious alteration in the substance dissolved.

## EVAPORATION

**95. Evaporation assisted by heat.**—Put some water in an evaporating-basin placed upon sand in a shallow iron saucer, known as a “sand-bath,” or upon a “water-bath” made by placing a shallow dish upon the top of a beaker of boiling water (Fig. 58). Warm it gently by means of a laboratory burner or spirit lamp. Notice that as the heating is continued the water in the basin gradually disappears or evaporates.



FIG. 58.—A simple water-bath.

Recall what you know of the drying of a wet road by evaporation.

**96. Cooling produced by evaporation.**—Sprinkle a few drops of (1) spirits of wine, (2) carbon bisulphide, (3) ether, on your hand in succession. Notice that the liquid soon disappears and its presence in the air can be detected by its smell. The rate at which the liquid evaporates is increased by waving the hand about.

**97. Difference between evaporation and boiling.**—Boil some water in a flask, and observe that, when boiling begins, bubbles of steam are formed everywhere throughout the liquid, and that the bubbles rise and burst at the top.

**Evaporation.**—If a saucer of water is left for a few days, it will disappear, or, as is generally said, dry up. The water can be made to disappear more quickly by gently heating it. When a solution containing salt or sugar is made to dry up in this

way, the salt or sugar does not disappear, but remains in the saucer. The name given to this process of turning a liquid into a vapour is *evaporation*. The solid left behind is spoken of as a *residue*.

**Everyday examples of evaporation.**—In summer it is not long before a road becomes dry again after having been well watered by a water-cart. Wet clothes hung upon a line and exposed to the sun and air soon become quite dry. Streams and rivers, ponds and lakes, are shallower during the hot days of summer than they are in other months of the year. It is a common practice to expose shallow vessels of water in rooms which are warmed during winter by coke- or gas-stoves. From

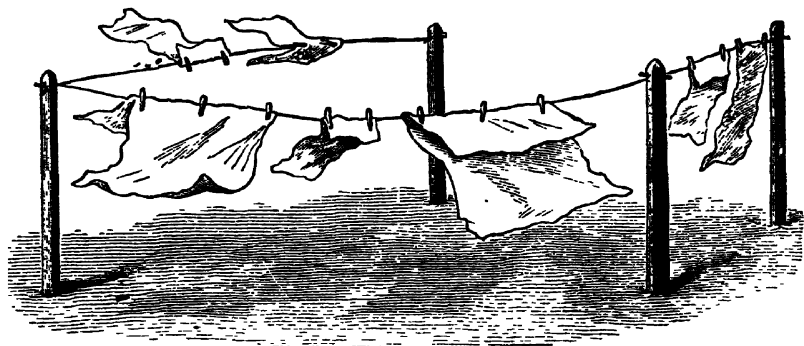


FIG. 59.—One way in which evaporation is utilised

time to time such vessels have to be refilled, for the water quickly disappears and exists in an invisible form throughout the air of the room. Nearly everybody has observed the same process taking place when sitting still after some violent exertion. The perspiration, which after such exercise stands in drops over the face and body, soon disappears by passing into the air as vapour, and leaves the skin dry. The cooling of the body which is also noticed at the same time will have to be explained a little later.

**Rate of evaporation.**—Some days are better for drying clothes than others. A good “drying day,” as the laundress calls it, is not only a warm day, but also one when, as well as this, there is a good breeze. Even when the day is warm, if

at the same time the air is still and damp, the wet clothes do not dry quickly. For the same reason a little spirit placed upon the hand dries up more quickly when the hand is moved about than when the hand is kept still.

It is well known that a sponge can absorb a good deal of water, but soon takes up as much as it can hold, when it is said to be *saturated*. In much the same way a given amount of air can only hold a certain quantity of the invisible vapour into which liquids are changed by warming them. If the air near a liquid is quite still, it soon has as much vapour as it can hold, and evaporation from the liquid ceases. If, on the other hand, the air is continually moving, and fresh quantities of air are constantly being brought near the liquid, evaporation never ceases so long as any liquid is left.

**Coolness caused by evaporation.**—When a liquid is changed into vapour a certain amount of heat is used up. It does not matter whether the liquid evaporates or boils; every gram of it requires a certain amount of heat before it becomes converted into vapour. In boiling, this heat is supplied by the flame or fire, and in evaporation it is taken from the objects in contact with the liquid. The faster the evaporation the more heat is absorbed in this way. When a liquid evaporates very rapidly, the cooling produced is very noticeable. So much heat may be absorbed in this way that, as illustrated by the experiment shown in Fig. 44, water can be frozen by the evaporation of ether in a vessel in contact with it.

**Quiet evaporation and boiling are not the same.**—It is very instructive to watch some vigorously boiling water in a flask or beaker, and compare it with the evaporation of a solution in an evaporating basin, where only a small flame, or no flame at all, is used. In the case of the solution being gently evaporated, all the vapour is formed *at the surface* of the liquid, and the process goes on quietly until no liquid is left in the basin.

In the case of vigorously boiling water, bubbles of vapour are formed *everywhere throughout the mass of the liquid*. The bubbles can be seen at the bottom and the sides, and they rise from every point to the surface, each as it escapes there making a little noise. The sounds of the bursting bubbles added together make up the “singing” or “rattling” which is heard when water is briskly boiling in a flask or other vessel.

## DISTILLATION

**98. Condensation of water vapour from the air.**—Put some pieces of ice into a test-tube, or other glass vessel, which is clean and dry on the outside. In a very few minutes the outer surface of the vessel will become covered with moisture, owing to the *condensation* of water vapour from the air.

If you cannot procure ice, put a little ether in a test-tube and make it evaporate quickly by blowing vigorously down a narrow tube on to the surface of the ether. As the air is cooled by the evaporation of the ether, the moisture in the air is given up and deposited on the outside of the test-tube in the form of minute drops; in other words, the water vapour in the air is condensed.

**99. Distillation of Water.**—Obtain a glass retort and put in some

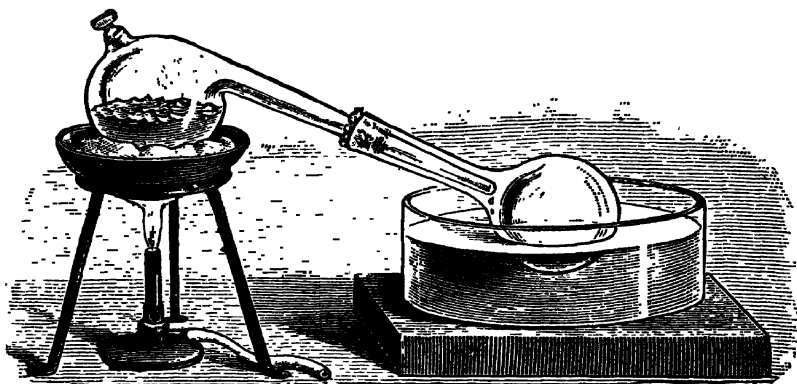


FIG. 60.— A simple means of purifying water by distillation.

ordinary tap-water coloured with ink, and in which some sand has been placed, and put it on a sand-bath, as shown in Fig. 60. Let the neck of the retort pass into the neck of a flask suitably supported over a basin of water. By means of a gas-burner boil the water in the retort, and keep the flask cool by continually pouring water on to it. Notice that the steam which passes over into the flask is condensed again into water, which is quite clear and tasteless.

If a condenser, like the one shown in Fig 61, is available, examine it, and use it to distil some dirty water in which salt has been dissolved. Taste the distilled water.

**100. Distillation of Beer.**—Take a test-tube, provided with a graduated piece of paper stuck inside along the length of the tube, as described in the experiment on p. 35, in explaining the principle of the lactometer. Load it with shot until it floats to a convenient mark on the scale in a sample of beer which has been allowed to get "flat." Indicate this mark in a suitable way. Now place some of the beer in

a distilling retort, as in the last experiment, and, as in the case of water, distil over sufficient liquid to test with the loaded test-tube. It will be found that the distilled liquid has a lower density than the beer; for the test-tube floats deeper in the liquid. Read the division of the scale on a level with the surface of the liquid, and so compare its density with that of beer.

**Condensation.**—By heating, a liquid is changed into vapour. This change may take place slowly and gently, as in evaporation; or quickly and vigorously, as in boiling. But by whichever process vapour is obtained, it can, by the reverse plan of cooling it, be reconverted into liquid. This change, from the state of vapour back again to the condition of liquid, is called *condensation*; the vapour is condensed to a liquid. Thus, a cold plate held close to the spout of a kettle from which steam is coming will cool the steam, condensing it into water, which will be seen trickling down the plate.

Most people have noticed the condensation of vapour taking place at some time or other on a cold day. For instance, if the doors and windows of a room are kept tightly closed, and there is a good fire burning, the inside of the window panes soon becomes covered with moisture, which, forming drops, trickles down the glass and collects on the window frame as liquid water. The water must evidently come from the air in the room.

The air outside a room is, in winter, much colder than that inside. This cools the glass of the windows very much, and consequently the air next to the cold surface itself becomes cooled, and then cannot hold so much vapour as when it is warm, and some of the vapour which can no longer be held by the air is changed into water. So that condensation is caused when air containing water vapour is cooled.

**Distillation.**—The change of liquids into vapours by heating them, and the condensation of the vapour into liquid by cooling it, is employed in an important process called *distillation*. This plan is frequently made use of for purifying water and other liquids. Perhaps the most useful application of distillation is to obtain fresh water for drinking purposes from sea-water or other water not fit to drink. Large ships, carrying as they often do more than a thousand people, cannot take enough fresh water on board for the needs, throughout a long voyage, of so

many persons. Instead of attempting this difficult task, it is the custom to change sea-water into fresh water by distillation. Ordinary water may be boiled in a flask or retort, the neck of which fits into a flask continually kept cold by pouring water upon it, as in Fig. 60. As the water boils it is converted into steam, and this steam is condensed again into water in the cooled flask. If some ink and salt were first added to the water in the retort, they would be left behind, and the water found in the flask would neither taste of the salt nor be coloured by the ink; it would be purified from these by distillation. So that

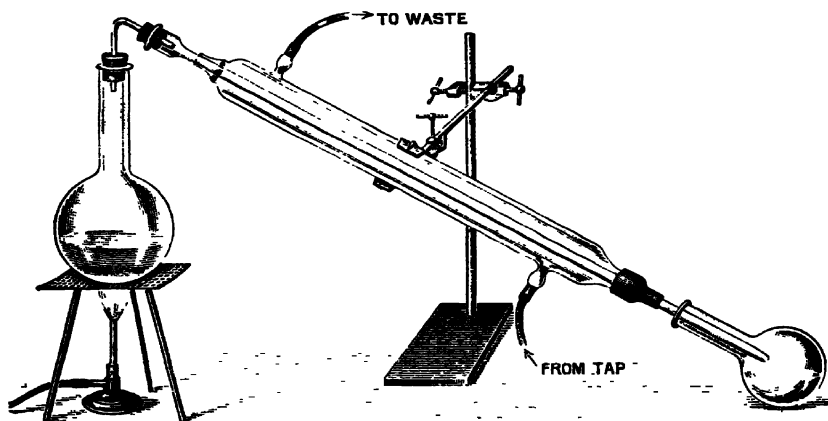


FIG. 61.—Distillation with a glass condenser.

all impurities which water contains are left behind in the retort, and the water obtained by condensing the vapour is pure.

**Distilling apparatus.**—A convenient apparatus for distilling a liquid is shown in Fig. 61. The water, or other liquid, to be distilled is placed in a retort which is put on a piece of wire gauze, or a sand-bath, and is then boiled. The retort is connected with an arrangement called a *condenser*. This consists, in the illustration, of a narrow glass tube surrounded by a wider one, having a short connecting tube at the side near each end. One of these short tubes is connected with the water-supply; the other carries the waste water to a drain or sink. In this way a stream of water can be made to circulate through the wider tube outside the inner narrow one, which it consequently keeps cool. A second flask is placed at the free end of the long

narrow tube to catch the pure water as it is condensed. When water is boiled in the retort, steam passes along the inner narrow tube, and is cooled and condensed by the water in the outer and larger tube, and the drops of water formed pass on into the second flask. The water in this flask is thus the condensed steam only ; for if the apparatus be properly constructed none of the water which flows around the inner tube can get into the receiving flask. The stills used for obtaining fresh water

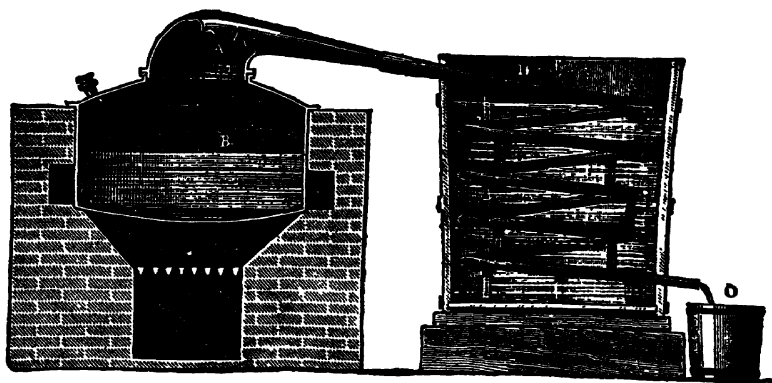


FIG. 62.—A large metal “still.” The vapour of the liquid boiling in B passes through the neck A and the spiral tube in D. The spiral is surrounded with cold water, so the vapour in it is condensed, and trickles in drops of liquid into O.

from sea-water are constructed on exactly the same plan. All that is wanted is a vessel to boil the water in, and an arrangement for condensing the steam and collecting the drops of water formed. An arrangement for distilling large quantities of water is shown in Fig. 62.

#### SUMMARY.

**Solution** is the process by which some substances, when placed in water or other liquids, disappear, and their particles spread through the entire liquid. A substance is said to be *soluble* in a liquid when it disappears in it, forming a solution. Sugar, salt, and soda are soluble in water.

A substance which will not dissolve in a liquid is said to be *insoluble*. Sand, camphor, and sulphur are insoluble in water. Substances insoluble in water often dissolve in other liquids. Camphor is soluble in spirits of wine ; sulphur dissolves in carbon bisulphide.

Insoluble substances may be spread throughout a liquid, that is, *held in suspension* in it. Suspended impurities can be got rid of by *filtering*.

Some *liquids mix with or dissolve in* other liquids. Whisky, brandy, and vinegar dissolve in water. Some liquids do not mix or are insoluble in one another. Oil, water, and mercury do not mix.

Some gases are soluble in liquids. Carbon dioxide and ammonia dissolve in water.

**Evaporation** is the process of changing a liquid into a vapour by heat. Water and other liquids are only converted into vapour by taking up heat—or, *evaporation causes cooling*. Some liquids evaporate more quickly than others at the same temperature.

Evaporation takes place more rapidly when the air above the liquid is in motion.

**Boiling** and quiet evaporation are not the same. In boiling, vapour is formed everywhere throughout the liquid; in quiet evaporation, only at the surface.

**Condensation** is the process by which vapours, such as steam, are reconverted into liquids by cooling.

**Distillation** consists in first converting a liquid into a vapour by boiling, and then reconvertng the vapour into liquid by condensation. Distillation is a means of getting *pure (distilled) water* from water having dissolved impurities in it. Rain water is really distilled water.

### EXERCISES ON CHAPTER VII.

1. Distinguish between soluble and insoluble substances. Which of the following are soluble in water? Sand, sugar, table-salt, camphor, sealing-wax, baking-powder, and coal.

2. Some solids which do not dissolve in water are soluble in other liquids. Give all the instances of this with which you are acquainted.

3. What class of impurities can be removed from water by filtering? Explain clearly how you would proceed to get clear water from muddy water.

4. Describe fully how pure water can be obtained from sea-water.

5. Do gases dissolve in water? If you think they do, name two or three soluble gases.

6. What exactly is meant by saying that some liquids do not mix? Mention examples of liquids which do not mix, and also of liquids which do mix.

7. Which substance dissolves more readily in water, salt or sugar? Explain fully how you would show what you say to be true by an experiment.

8. How is it that clothes placed on a line to dry are ready to be taken in sooner on some days than others? What sort of day does the laundress consider a good drying day?

9. In both quiet evaporation and boiling a liquid is converted into vapour. What differences are there between the two processes?

10. The cooling produced by evaporation is sometimes enough to cause water to freeze. Describe an experiment to demonstrate this fact.



11. What is meant by condensation? Name two or three familiar instances of the condensation of water vapour.

12. Explain the terms : Distillation, "held in suspension," filtering, and insoluble.

13. You are given a mixture of sand and salt. Explain clearly how you would separate the two substances.

14. Gunpowder is a mixture of nitre, sulphur, and charcoal. If you were given some gunpowder, how could you get specimens of each of the three things from it?

## CHAPTER VIII

### THE ATMOSPHERE

#### PRESSURE OF THE ATMOSPHERE

**101. "Empty" vessels contain air.**—Try to force an empty bottle, held upright with its mouth downwards, into a vessel of water. When you leave go, the bottle jumps up again. There is something in it which acts like a spring. Tilt an empty bottle, held mouth downwards, in a trough of water, and notice the bubbles of air which pass up as the water enters the bottle.

**102. Forcing air from an "empty" bottle.**—Take a funnel with a narrow neck, and fit it firmly into a bottle by means of an india-rubber stopper with two holes in it. Through the second hole pass a short piece of glass tubing bent at right angles (Fig. 63).

Place a finger over the open end of the tube, pour water into the funnel, and notice that, so long as you keep your finger upon the end of the glass tube, the water is prevented from getting into the bottle by something—air—already there (Fig. 63). Take your finger away. The water now runs into the bottle, and air escapes from the tube. The escaping air may be felt, or its effect upon a lighted match may be seen.

**103. Air has weight.**—(a) The following experiment needs to be carefully done with a good balance in order to be successful. Obtain two large flasks. Fit one with a closely-fitting india-rubber stopper, having a hole in it through which a short piece of glass tubing passes. Upon an india-rubber tube attached to the glass tube a clip is fastened. Hang the flasks over the two pans of the balance, the fitted one having the stopper and tubing with it, and counterpoise them.



FIG. 63.—The bottle only contains air, but the water will not run in until the finger is taken from the tube, and then the air can be felt coming out.

Now insert a short glass mouthpiece in the india-rubber, open the clip, and suck air out of the stoppered flask, without touching the flask. Fasten the clip while you are sucking out the air. Take out the mouthpiece, and you will find that the flasks no longer counterpoise, the one from which air has been withdrawn being lighter than before. Admit air by opening the clip, and it will be found that the flasks again counterpoise one another.

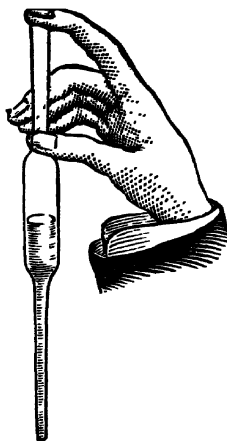


FIG. 64.—Use of a pipette.

(b) If a metal bottle with a tap is available, together with an air-pump, this experiment of showing that air has weight may be repeated with advantage. It is impossible to suck out all the air; but, by means of the pump and a metal bottle, there is no difficulty in demonstrating the fact with ease.

(c) A third way to show that air has weight is as follows:—Into a flask fitted with an india-rubber stopper perforated by a glass tube, as described under (a) of this section, pour a little water and counterpoise the flask, water, stopper, and clip, upon a balance. Now place the flask upon a sand-bath, open the clip, and boil the water. So soon as steam issues from the tube, close the clip, remove the flask, and allow it to cool. Place it upon the scale pan, and by trial satisfy yourself that the flask is much lighter than before.

The steam drove the air out of the flask, and the flask is consequently lighter. It must be remembered, however, that part of the loss of weight is due to the escape of a little steam.

**104. Warm air is lighter than colder air.**—Round the neck of the flask, fitted with a stopper, like that used in the last experiment, stick a piece of gummed paper to show how far the stopper goes in. Push the stopper down into the flask to the mark, and find the weight of the closed flask; at the same time read the temperature of the air.

Now plunge the flask, after taking out the stopper, into hot water in a large basin, and by means of a thermometer find the temperature of the water in the basin after the flask has been in it five minutes, and at the moment of taking the temperature insert the stopper.

Push the stopper down to the mark, dry the flask, and place it on one side to cool. Weigh the flask when cold. It will be lighter than before.

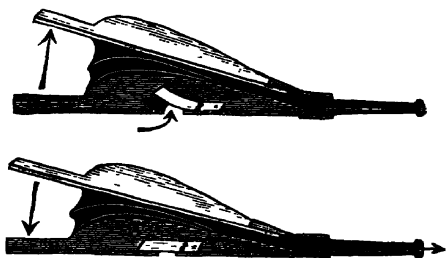


FIG. 65.—Section of a pair of bellows to show the action of the valve.

**105. Experiments depending upon air pressure.**—(a) *Pipette.*—Dip a long tube—a pipette will do—into water. Place your finger over the top and lift the tube out of the water. Notice that the water does not run out of the tube although the bottom is open (Fig. 64).

(b) *Suction dependent upon air pressure.*—Fit a one-holed stopper into a flask of water. Push a piece of glass tubing through the stopper. Try to suck up the water (Fig. 66). You cannot, unless you loosen the stopper so as to let the pressure of the air force the water up the tube.

(c) *Action of bellows.*—Procure a pair of bellows. Notice that the valve at the bottom only opens inwards. Open the bellows, and observe that the valve is pushed

up a little as the air enters (Fig. 65). Close the bellows; the valve is pushed down; and, as the air cannot escape any other way, it is forced through the nozzle.

(d) *Upward pressure of air.*—Take a tumbler or cylinder with ground edges and completely fill it with water. Place a piece of card or stout writing paper across the top and invert the vessel. If the air has been carefully kept from entering the tumbler, the

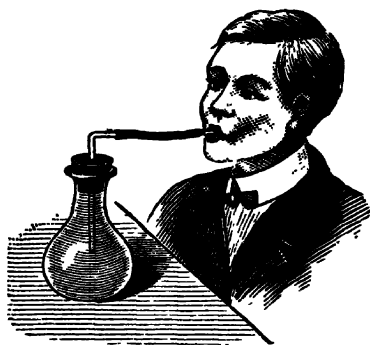


FIG. 66.—The water in the flask cannot be sucked out while the cork is tight.

water does not run out (Fig. 68). Think what keeps the paper in its place.

**106. Principle of the barometer.**—Tie a short piece of thick india-rubber tubing upon the open end of a barometer tube. Tie the free end of the tubing to a glass tube about six inches long, open at both ends. Rest the barometer tube with its closed end downwards, and pour mercury into it (being careful to remove all air bubbles) until the liquid reaches the short tube. Then fix the arrangement upright as in Fig. 67.

The air pressing upon the surface of the mercury in the short open arm of the U-tube balances a long column of mercury in the closed arm.



FIG. 67.—A convenient form of barometer.

**There is air all round us.**—Very few observations are required to convince any one that there is air around all objects

on the earth's surface. It can be felt by any one who has to walk against the wind. The effects of air in motion can be seen by watching the branches and leaves of trees moving about in a high wind ; and air in motion is usefully employed in driving windmills, and in blowing sailing ships across the ocean. At times the air moves with sufficient force to blow down houses, uproot trees, and do other damage. Even on the calmest day the air can be felt by waving the hand about or by using a fan.

What is usually called an "empty" bottle is really one full of air. If the bottle is dipped under water and tilted, the air can at once be seen bubbling through the water so as to make room for water to enter the bottle. Or, if water is allowed to run into the bottle through a funnel passing through a cork, the air can be felt as it escapes from a hole in the cork. A similar and more familiar instance of the presence of air is obtained by using a pair of bellows.

**Air has weight.**—Like every other thing about the existence of which we are aware by means of our senses, the air has weight. Substances are sometimes spoken of as being as light as air, but there are many gases known which are much lighter than air. A good example of a very light gas is the one called coal-gas, which is burnt in our houses. But compared with water, for example, air is very light. A cubic foot of air only has a weight of a little more than an ounce, whereas a cubic foot of water weighs a thousand ounces. Water is, in fact, about eight hundred times heavier than an equal bulk of air. To find the weight of the air in a flask, the flask must be first weighed full of air, and then when the air has been removed, which can be done with the help of an air-pump. The difference between the two weighings shows the weight of the amount of air necessary to fill the flask at the temperature of the room at the time of the experiment.

**As air is warmed it becomes lighter.**—It has already been learnt that air expands as it is warmed. That is to say, a given weight of air occupies a larger and larger volume as it gets warmer and warmer. This statement must be thoroughly understood ; supposing a cubic foot of air at the temperature of the room weighs one ounce ; if that air is heated to, say, the temperature of boiling water, being allowed to expand freely

the whole time, it will at the higher temperature occupy a much larger volume. Suppose, for the sake of example, that it becomes at the temperature of boiling water a cubic foot and a third. There is the same weight of air in the small volume at the lower temperature and the larger volume at the higher temperature ; that is to say, a cubic foot of air at the higher temperature will weigh considerably less than the cubic foot of air at the lower temperature.

This fact is of the greatest importance in ventilation. The air of dwelling-rooms becomes warm and impure at the same time ; the impure air which has to be got rid of is found consequently at the highest parts of the room, and generally some means for its escape into the outer air is arranged near the ceiling of such a room.

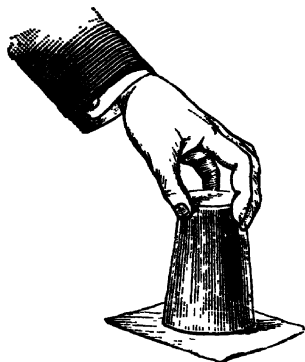


FIG. 68.—The paper does not fall off though the glass is full of water.

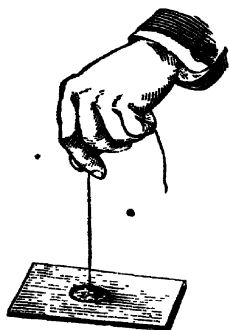


FIG. 69 —The sucker cannot be pulled off easily, because the air is pressing on the top of it, but not on the lower side.

**Air exerts pressure.** — Everything which has weight can exert pressure. The pressure depends first of all upon the density of a thing. For instance, if a piece of iron is carried upon the shoulder, it presses down more than a piece of wood of the same size. But the pressure also depends, of course, upon the amount of material which is borne. Thus, a sack of wool borne upon the head would exert more pressure than an iron nail. It is easy to understand, then, that though the air is so light compared with other substances, a large quantity of it would be very heavy, and would

exert a very great pressure.

Now, the air extends upwards from the surface of the earth for many miles ; and in consequence of this it presses very heavily upon everything. The reason why a leathern sucker is difficult to pull off an object upon which it has been placed is

that the air is pressing upon the outside of the sucker but not upon the inside (Fig. 69).

When a liquid is drunk through a tube, the vessel containing the liquid must be open to the air or else no liquid can be obtained. This can be seen by examining a baby's feeding-bottle. If the stopper of the bottle is screwed in tightly, and there is no open hole in it, the baby cannot get his food, however hard he sucks.

The pressure of air is not only downwards, but upwards and sideways; in fact, in all directions. If the pressure were only exerted downwards, then a sucker could easily be pulled off an object on which it is fixed by turning the object upside down or sideways. But the sucker cannot be pulled off any more easily whatever way it is turned.

Though our bodies are pressed upon by the whole weight of the air above us, we do not feel it. Why is this? The lungs, which fill up a large part of our chest space, are filled out with air, and this air is in free communication with the atmosphere through our throat and nostrils. The result is that the air in the lungs presses them outwards from the inside just as much as the atmosphere presses the chest inwards from the outside, and so we feel no inconvenience.

**The barometer and its use.**—It is very important to be able to determine how much the air presses upon things at the earth's surface. The instrument shown in Fig. 70 enables this to be done. The top of the long tube is sealed, so that the air cannot press upon the mercury in it, but the small tube is open, and the air can therefore press upon the mercury in it.

In an instrument like that shown in Fig. 70, and known as a *barometer*, a column of mercury is supported by some means which is not at first plain. If there were nothing pressing upon the mercury in the short tube, the mercury would sink to the same level in both tubes, for liquids, when free to move, always

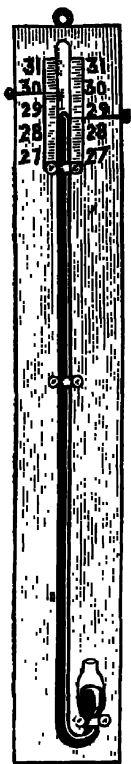


FIG. 70.—A common form of "siphon" barometer. The pressure of the air on the mercury in the small tube is able to support the column of mercury in the long tube.

find their own level. If a hole were made in the closed end of the tube, this would immediately happen. \*There should be no difficulty in understanding that the long column of mercury is kept in its position by the downward force of the weight of the atmosphere pressing upon the surface of the mercury in the short open tube. The weight of the column of mercury, and the weight of a column of the atmosphere of the same size through, or of the same area, is exactly the same. Both the column of mercury and the column of air must be reckoned from the level of the mercury in the short stem of the barometer shown in Fig. 70—the mercury column to the top in the long tube; the air to its upper limit, which is at a great distance from the surface of the earth.

For an instrument of this kind to be accurate, great care has to be taken that no air enters the space at the top of the long tube. If air does enter, it will press upon the surface of the mercury in the long tube, and the height of the mercury will be less than thirty inches. In such a case, instead of measuring the whole pressure of the atmosphere, what we should really be measuring would be the difference between the pressure of the whole atmosphere and that of the air enclosed in the longer tube. In a properly constructed barometer, therefore, there is nothing above the mercury in the longer tube except a little mercury vapour.

### MOISTURE IN THE AIR

STRONG SULPHURIC ACID SHOULD BE HANDLED WITH GREAT CARE.

**107. The water in air.**—Repeat the experiments on condensation on p. 85.

**108. Moisture taken from the air by calcium chloride.**—Place some lumps of dry white calcium chloride in a basin, and find the weight of the basin and its contents. Expose the calcium chloride to the air for two or three hours. Notice the solid has disappeared, and a liquid is present in its place. Find the weight of the basin and liquid; you will find it greater than before. The increase shows the weight of the water taken out of the atmosphere by the calcium chloride.

**109. Moisture taken from the air by sulphuric acid.**—Repeat the last experiment, substituting a small wide-mouthed bottle for the basin, and strong sulphuric acid for the calcium chloride. Leave the acid exposed for a day, and, as before, notice the increase of weight. Weigh again after a second day's exposure, and notice the further increase of weight.

**110. Water in the air due to evaporation.**—(a) Put some water in



an evaporating-basin. Determine the weight of the basin and water. Place the basin, with the water in it, in an exposed position free from dust. Determine its weight every day for a week by weighing. Notice that the weight decreases daily.

(b) Repeat the experiment, using a roll of wet flannel instead of the basin of water.

**There is always water vapour in the air.**—Since evaporation is always going on from every surface of water, whether it be that of sea, lake, or river, there must always be some water vapour present in the air. The amount on hot days is greater than on cold ones, because the warmer the air the larger the amount of moisture it can take up. For this reason clothes usually dry more quickly on hot than on cold days. When describing

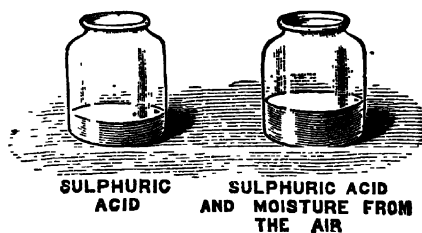


FIG. 71.—When sulphuric acid is exposed to the air for a few days, it absorbs so much moisture that a rise of level is produced.

condensation, some of the ways in which the presence of this water vapour can be shown were explained. Whenever a cold surface comes into contact with warm air, the water vapour which the air contains is deposited in the form of visible moisture.

**Some substances take water vapour out of the air.**—If the salt on the dinner-table is examined from time to time, it is noticed that on some days it is moist, on others quite dry. Some persons hang up a piece of seaweed outside their houses, and observe that while on some days it is stiff and dry, on others it is moist and limp. The difference in the dampness of the salt or the seaweed on different days depends upon the amount of moisture obtained by them from the air.

Another simple experiment may be made to show the power some substances have of absorbing moisture from the air; and, of course, it also proves that there is really water vapour in the atmosphere. Calcium chloride and strong sulphuric acid (oil of vitriol) both have this power of taking up moisture. Calcium chloride, in the form in which it is generally used in the laboratory, is a white solid. If a lump of it is placed in an

evaporating-basin, and left freely exposed to the air for an hour or two, it is found on examining the contents of the basin after this time that, in the place of the dry white solid lump taken from the bottle, there is some liquid only to be found. This liquid is the water which has been absorbed from the air and has dissolved the lump of calcium chloride.

If the basin and dry calcium chloride are carefully weighed before exposing to the air, and the basin and liquid are again weighed after all the solid has disappeared, it will be found that the liquid weighs more than the calcium chloride which was used. In fact, the difference in the two weighings shows the weight of the water which the solid has taken out of the air.

**The water vapour in the air sometimes becomes visible.**—One instance of this has been already mentioned. When a person breathes on a cold frosty day, the moisture in his breath becomes condensed and visible as a cloud of mist in the air in front of his mouth.

Whenever the air is cooled sufficiently, there is a condensation of water vapour, and the drops of water thus formed build up one of the following forms of condensed water vapour :

- |                     |                                 |
|---------------------|---------------------------------|
| (a) Mists and fogs. | (d) Snow, hail, and hoar-frost. |
| (b) Clouds.         | (e) Dew.                        |
| (c) Rain.           |                                 |

Mists, fogs, and clouds are very much alike, and all of them consist of small droplets of water condensed from the air. When a number of these droplets run together, on account of further cooling, drops of rain are produced. When the temperature is below the freezing point of water, hoar-frost and snow are formed instead of mist and cloud from the moisture in the air. You have probably noticed frost patterns upon window-panes in the winter. The patterns consist of frozen moisture from the air.

**Effects of the water vapour in the air.**—It is important for some purposes that the exact amount of water vapour in the air should be known. For instance, the "dryness" or "wetness" of the air has a very great influence upon health. Some persons are unable to keep in good health in a damp atmosphere. For this, and other reasons, it is often necessary to find out exactly how much water vapour the air contains. But it must not be supposed from this statement that

the actual quantity of such moisture in the air decides whether the air is damp or dry. Air which *feels* damp may really only contain a small amount of vapour, and perhaps very much less than air which feels dry. *Air feels dry when it contains considerably less water vapour than it is capable of holding. Damp air has as much, or nearly as much, water vapour as it can possibly take up.*

## AMOUNT OF MOISTURE IN AIR

**111. Determination of proportion of moisture in air.**—Obtain a piece of glass tubing about  $\frac{1}{2}$  inch diameter and about 10 inches long. Fit a good cork into each end, and push through each a piece of narrow glass

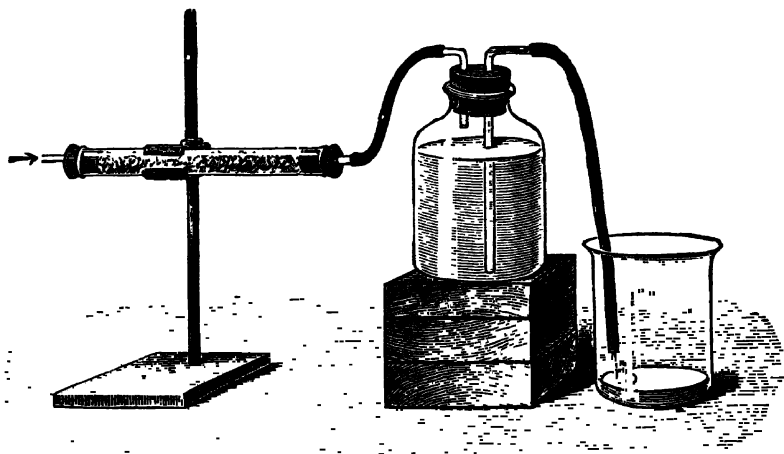


FIG. 72 — Air is being drawn through the horizontal tube, which contains calcium chloride to absorb the moisture in the air.

tubing. Fill the tube with pieces of calcium chloride, and weigh the whole. Support the tube in the clamp of a retort-stand, and connect one end by means of an india-rubber tube to an aspirator. Allow the water slowly to run out of the aspirator, when air will pass over the calcium chloride in the tube. Measure the water which runs from the aspirator, and so determine the volume of air which passes through the tube. At the end of the experiment, again weigh the tube and calcium chloride, and so find the increase in weight.

**112. Rise of water in salt and sugar.**—Stand a lump of sugar and a lump of salt in a saucer containing a little ink and water. Notice that the liquid rises up the solids.

**113. Rise of water in threads.**—Place the corner of a dry handker-

chief in a glass of water. Notice the rise of the water along the threads of the handkerchief, which in time becomes wet over a large part.

**114. Construction of a hygrometer.**—Obtain two precisely similar unmounted thermometers and suspend them from a suitable support, such as the ring of a retort-stand (Fig. 73). Cover the bulb of one of the thermometers with a square of muslin tied up round it to form a bag. The muslin is best tied just above the bulb by a piece of thread. To this piece of thread attach several other long pieces of thread, and let them dip into a glass of water. When the muslin has got thoroughly damp, compare the readings of the two thermometers. Notice that the temperature recorded by the one with the wet muslin round it is lower than that shown by the other.

The two thermometers used in this way form what is known as a *hygrometer*, or a *wet- and dry-bulb thermometer*.

**The wet- and dry-bulb thermometers.**—Instruments which are used to measure the amount of water vapour in the air are called *hygrometers*, and the apparatus shown in Fig. 72 is a *chemical hygrometer*. There are many different kinds of hygrometers, but only one other form need be described here.

A kind of hygrometer which is in very common use is shown in Fig. 74. It consists of two exactly similar thermometers which are usually fixed side by side on a suitable frame. One thermometer, however, has its bulb covered with a little muslin bag, from which hang several cotton threads. These threads dip into a vessel of water. This is called the *wet-bulb thermometer*. The other, which only differs from the wet-bulb thermometer in having no wet muslin round it, is the *dry-bulb thermometer*.

This hygrometer depends for its use upon two facts which have already been stated. The first is, that when liquid water is changed into water vapour it must take up heat in doing so. The other fact is, that the quantity of water vapour which air can take up depends upon the amount it already contains.

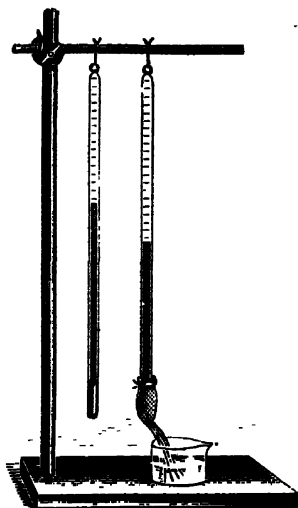


FIG. 73.—Wet- and dry-bulb thermometers.

Now let us see how these facts are utilised in the instrument. When the cotton threads hanging from the muslin bag, surrounding the bulb of one thermometer, dip into the vessel of water, they soon become wet all the way up to the bag. After a moment or two the bag itself gets wet, and so long as the threads are hanging in the water the bag remains wet.

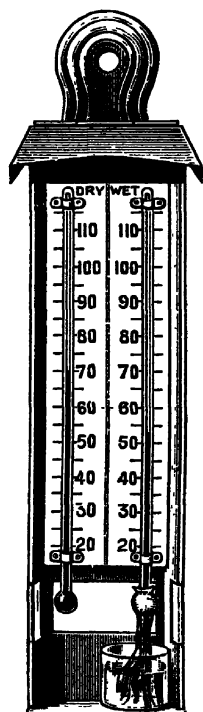


FIG. 74.—A hygrometer for showing whether the air is dry or moist.

But, as you also know, evaporation will begin to take place from the muslin bag so soon as it is wet. The heat necessary for the change of the water into water vapour is taken from the bulb of the thermometer inside the bag, and the bulb and quicksilver in it consequently get cooler. This causes the height of the quicksilver in the stem of the thermometer to fall ; in other words, the temperature recorded by the thermometer is lower. The more water which is changed into vapour at the surface of the bag, the more heat is taken from the bulb of the thermometer, and the lower is the temperature shown. This goes on until the air round the muslin bag can take up no more moisture.

If, therefore, the air is dry to start with, the water on the muslin bag evaporates quickly, and the temperature of the wet-bulb thermometer consequently is much lower than that of the other. But the dry-bulb thermometer gives the temperature of the air, and so we see that the drier the air to start with, the greater is the difference in the readings of the two thermometers. It is the difference in the readings of these thermometers which is noticed and enables the observer to tell whether the air is dry or damp.

Hygrometers of this kind are frequently kept in cotton mills and other factories to indicate whether the air is in a proper condition of dampness or dryness.

## SUMMARY.

The **air around us** can be felt when the wind is blowing ; the results of its movement can be seen in the swaying branches of trees, or when trees are uprooted by the wind.

The **weight of air** can be determined by weighing a tightly closed bottle full of air, and again after all the air has been pumped out.

The **pressure of the air** is due to its weight ; its existence is demonstrated by using either a sucker or a squirt.

When **air is removed** by suction from the inside of a tube which is dipping in water, the pressure of the air forces the liquid up the tube.

A **barometer** is an instrument for measuring the pressure of the air. The principle of its action is that a column of mercury in a tube containing no air is balanced by the pressure of the atmosphere outside the tube.

**Warm air** is lighter, bulk for bulk, than cold air.

There is always **water vapour in the air**, due to the continual evaporation going on from every exposed surface of water.

**Some substances**, like calcium chloride and sulphuric acid, **absorb moisture** from the air. This fact makes it possible to ascertain *how much* water vapour has been taken out of a quantity of air.

*Fogs, clouds, rain, snow, hail, &c.*, are formed by the condensation of the water vapour of the atmosphere.

**Hygrometers** are instruments for measuring the amount of water vapour in the air. The *chemical hygrometer* depends upon the absorptive power of calcium chloride or sulphuric acid for water vapour. The *wet- and dry-bulb thermometer* consists of two similar thermometers. The bulb of one is covered with a muslin bag from which hang threads that dip into water. The muslin thus becomes wet, and the evaporation cools the thermometer bulb in the bag, lowering the reading of this thermometer. The difference in the readings of the thermometers is great when the air is dry, and small when the air has much water vapour in it.

## EXERCISES ON CHAPTER VIII.

1. Why cannot a baby get his food from a feeding-bottle, if the stopper contains no hole and is tightly screwed down ?
2. Describe an experiment to demonstrate that air has weight. What is about the weight of sixteen cubic feet of air ?
3. Describe and draw some common instrument that depends for its efficacy on the pressure of the air.
4. What is a barometer used for ? Explain how you would make a simple barometer.
5. What would happen if a hole were made in the top of a barometer ? Explain clearly how the reading of a barometer would be affected, if, during its construction, some air got into the closed tube above the mercury.

6. Describe an experiment to prove that warm air is lighter, bulk for bulk, than cold air.

7. Some strong oil of vitriol is left exposed to the air of a room, and, after a time, it is found that the glass containing the acid has more liquid in it than at first. Why is this?

8. How is it that the windows of a railway carriage get covered with moisture if they are closed for some time during a journey on a winter's day?

9. Give a short history of the frost figures often found upon one's bedroom windows on getting up on a winter's morning.

10. How is it that the air we breathe sometimes feels dry and sometimes damp?

11. What is a hygrometer? Explain as fully as you can the wet- and dry-bulb thermometer, and state what it is used to measure.

12. Omnibus horses sometimes give off from their bodies quite a cloud of steam. When should you expect to see this happen, and what is the cause of it?

13. Explain the trick of keeping the water in a full tumbler held upside down, by means of a sheet of stout paper.

## CHAPTER IX

### BURNING AND RUSTING

#### RUSTING OF IRON

**115. Increase of weight when magnesium is burnt.**—Weigh a crucible and its lid with a piece of magnesium, which, folded lightly, is placed in the crucible. Heat it strongly in a burner, taking care to let no fumes escape (Fig. 75). To do this, keep on the lid, and only raise it a little when the flame is removed. The magnesium is seen to burn brightly in places, but, if care is taken, no fumes are lost. When finished, the whole mass should be in the form of a white powder. Allow to cool, and weigh the crucible with the lid and powder. Subtract the weight of the crucible and lid to find the weight of the powder. The powder will almost certainly be found to weigh more than the original piece of metal.

**116. Increase of weight when iron rusts.**—Carefully weigh a watch-glass with some iron filings or tacks in it. Because iron rusts best when damp, add a few drops of water to the iron in the watch-glass, and allow it to stand for a day or two. At the end of this time warm the watch-glass gently, so as to evaporate any water left. When the rusty iron is *quite dry*, weigh the watch-glass and its contents again. Its weight will be found to be more after the rusting has taken place; evidently, therefore, the iron in getting rusty has gained in weight.

**117. Air absorbed during rusting.**—Place some iron filings in a muslin bag and tie the bag to a piece of glass rod. Moisten well, and place it in a bottle of air inverted over water. If necessary, put something on the bottle to keep it upright (Fig. 76). Examine after a few days. It

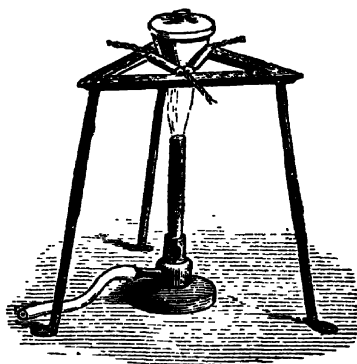


FIG. 75.—How to heat magnesium powder or ribbon in a crucible.



will be seen that the water has risen in the bottle, showing that some part of the atmosphere has been abstracted by the iron in rusting.

**118. Alteration of air when iron rusts in it.**—Place your hand or a card tightly under the mouth of the jar so as to allow no water to escape; set the jar upright, and plunge a burning taper into it. Notice that the flame is extinguished; do not throw away the water.

**Effects of heating metals in air.**—Some of the effects of heat upon substances have already been observed and described. It has been seen that, by heat, solids may be melted into

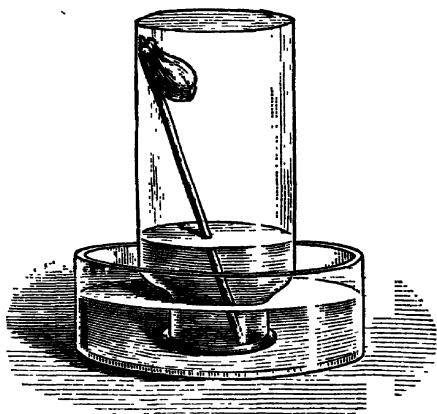


FIG. 76.—As the iron filings in the muslin bag become rusty, they use up part of the air in the bottle.

liquids, and liquids converted into vapours. We now come to another kind of effect produced when substances are strongly heated in the air. If a piece of platinum wire or platinum foil is held in the smokeless flame of a laboratory burner, it becomes red-hot; but when it is taken out of the flame it quickly resumes its ordinary colour, and no change can be seen to have taken place. But, if a piece of magnesium wire is held in a flame, it catches fire and burns away bril-

liantly, leaving a white ash. Platinum and magnesium thus behave very differently when heated. Many metals, like iron, copper, and lead, become covered with a film or tarnish when they are heated in the air, but if they are heated in closed tubes without air they do not change in this way. It thus seems that the tarnish is due to the absorption of something from the air.

But if something is taken from the air when a metal tarnishes, or when a metal like magnesium burns, the tarnished metal, or the ashes of the magnesium, should weigh more than the original substance. Observations show that this is actually the case. As air seems responsible for the effects described, it is evidently well worth investigation.

**Chemical properties of air.**—It is necessary carefully to

consider the changes which different substances undergo when exposed to the atmosphere. It is best to begin with those cases which appear to be simplest. Every one has noticed that when iron is exposed to damp air it becomes rusty. Does the iron lose or give up something when it rusts? Or, does it, on the contrary, take up or gain something? These questions can be best answered by properly arranged experiments.

**Iron gains in weight during rusting.**—If a known weight of iron is allowed to rust by contact with damp air, you can easily show by weighing it after the rusting has taken place that it has increased in weight. The result of this experiment is very important. If the weighings are carefully made, the iron is always found to gain in weight when it rusts.

**What causes the gain in weight?**—The substance causing the increase of weight, when damp iron filings rust, could come from the water, or moisture, on the iron, or from the air. If the iron is allowed to rust in a closed space, and the experiment is so arranged that if anything is taken from the air the loss can be detected, it can be decided whether the water or the air causes the rusting. Fig. 76 shows a convenient way of doing this. Some iron filings are placed loosely in a muslin bag, and the bag is tied to a piece of glass rod. The bag of filings is well moistened, and arranged in a bottle of air inverted over water in a basin, in the manner shown by the illustration. The apparatus is then left undisturbed for a day or two. When it is examined after this time, the water is seen to have risen in the bottle. Why is this? It is quite clear that there is less air in the bottle now than there was before the iron became rusty. Some part of the air has, therefore, been used by the iron as it rusted, and this part of the air has joined with the iron to help to make the rust on the outside of it.

**Air as well as iron undergoes change.**—When iron rusts, the change which it has undergone is visible. No difference can, however, be seen between the character of the air left in a bottle in which iron has rusted and ordinary air. But there is a very great difference. A flame is extinguished by the gas left in a bottle in which iron has rusted, hence the gas cannot be ordinary air, for a taper will burn quite easily in air. But before the rusting of the iron took place in it, the air *was* ordinary air. Hence, it is clear that the rusting of the iron is

accompanied by a change in the character of the air in the bottle. It is reasonable to suppose that the gas which disappears is concerned with the formation of iron rust, and this will be shown later to be actually the case. When iron rusts, it takes out of the air that part of it which helps burning, and moreover, the iron and the part of the air concerned in burning combine together to form iron rust. The part of the air left in the bottle will not let things burn in it. It should be very carefully remembered that : *Iron in rusting gains in weight, taking some material from the air, and this material is the part of the air which causes substances to burn in it.*

### CHANGE IN THE AIR CAUSED BY RUSTING

**119. Volume of air used up in rusting.**—Measure in a graduated vessel the quantity of water in the bottle from the experiment in the last section. This is equal to the quantity of gas which has been used up and has joined with the iron. Also measure the quantity of water the bottle holds. This gives us the volume of air the bottle originally held. What fraction of the total volume of air enclosed has been used up?

**120. Change produced in air by iron rusting in it.**—Repeat the experiment of allowing iron to rust in an enclosed volume of air. After the iron has been left for a day or two, and there is no further rise of the water, mark the level of the water in the jar by a narrow strip of gummed paper on the outside. Carefully introduce another muslin bag of iron which is not rusty. This can be done by using a large enough basin of water, and pushing the bag through the water, being careful to allow no more air to get into the bottle. Examine the bottle after another day or two. There is no further rise in the level of the water, and the iron is not rusty. Evidently the gas which is left will not allow more iron to rust in it, though it is colourless and transparent, like ordinary air.

**What proportion of the air is taken by iron in rusting?**—Only a certain proportion of the enclosed air is taken out of it when iron rusts. Suppose some damp iron filings are allowed to rust in an enclosed amount of air, contained in a bottle inverted over a basin of water. The amount of water which rises into the bottle can be measured by means of a graduated vessel. A moment's thought will tell you that, as this water gradually takes the place of the material which the iron uses out of the air, its volume must be the same as the volume of the material so taken out of the air. The amount of water the bottle holds when full can be easily found, and the

result shows the volume of air in the bottle to begin with. If observations of this kind are made, it will be found that, when the bottle is full, it has five times more water in it than it has after the iron has rusted. Even if the experiment is repeated several times with bottles of different sizes, the result is always the same. It is always found that *one-fifth of the volume of enclosed air is used up by iron in rusting.*

**Chemical composition of the air.**—That part of the air which helps substances to burn, and is taken out of the air by iron in rusting, may be called the *active part* of the air. That part which is left by the iron, and will not allow a taper or

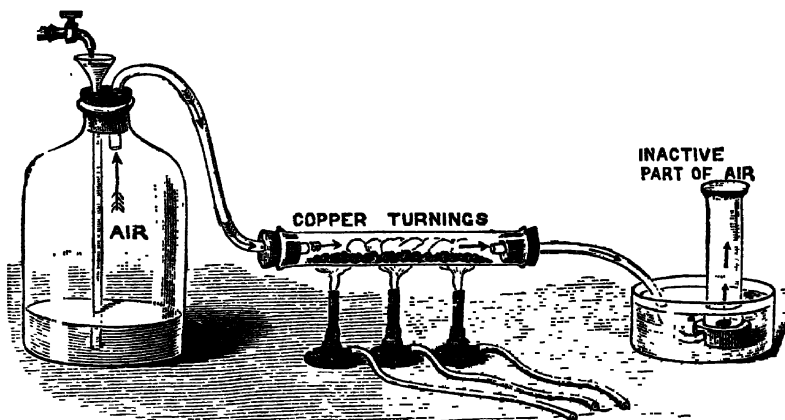


FIG. 77. When air passes over hot copper, it is deprived of its active part, and the inactive part may be collected as shown.

candle to burn in it, may be called the *inactive part*. The observations just described show that air is made up, or composed, of 1 volume of the active part to 4 volumes of the inactive part, in every 5 volumes. In other words, in 100 pints of air there are 20 pints which will unite with iron to make iron rust, or which will assist a candle to burn, and 80 pints of the inactive part which will not assist burning.

**Other metals combine with the active part of air.**—When copper is heated in air, it gradually blackens and increases in weight. When hot, copper has the power of combining with the active part of the air in just the same way as the iron does gradually when cold, and it is reasonable to conclude

that the black substance formed is copper rust, though it is not generally known by that name.

That copper only combines with the active part, and leave the inactive part, can be shown by a suitable experiment, though not so easily as in the case of iron. Some copper turnings are placed in a hard glass tube like that shown in Fig. 77, one end of which is connected with an aspirator full of air, and the other by means of a well-fitting cork with a tube which dips under water in a trough. Then a bottleful of water is inverted in the trough of water exactly over the end of the small tube, which dips into it, and is connected with the hard glass tube containing the copper. The copper is heated strongly, and air is forced over it by making water take the place of the air in the aspirator. As the air passes over the heated copper, the active part of the air joins with the copper to form the black copper rust, and the inactive part passes on alone into the inverted bottle in the trough. That the inactive part collects in the bottle is indicated, though not proved, by the fact that it puts out the flame of a taper. If this gas is, in the same way, passed over some

more heated copper, it has no effect on it; the copper does not blacken. Moreover, if the amount of air which has come out of the aspirator is measured, and also the amount collected in the bottle, it is found that in passing over the copper the air loses one-fifth of its volume.

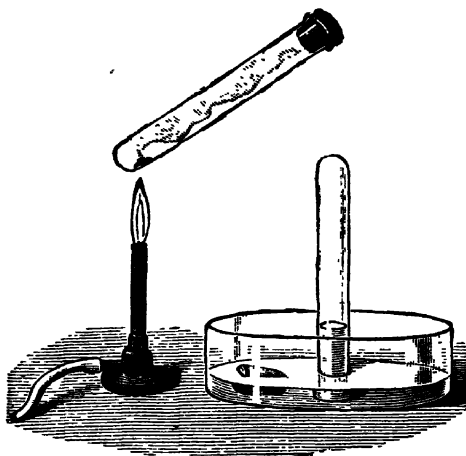


FIG. 78.—After the phosphorus has burned in the corked test-tube, the test-tube is uncorked under water, and water enters to take the place of the air used up.

## BURNING OF PHOSPHORUS

READ THE CAUTION ON  
P. III.

**121. The burning of phosphorus.**—Place a little phosphorus upon a slate,

tile, or an old saucer. Apply a light to it. It catches fire and burns brightly. As it burns, dense white clouds are formed.

**122. Volume of air used up when phosphorus burns.**—Place a little red (or yellow) phosphorus in a test-tube fitted with a good cork. Fix the cork firmly in the test-tube. Hold the test-tube slantingly, by means of a test-tube holder, over a flame for a second or two, so as to heat the phosphorus and make it burn. When it will burn no longer, take away the test-tube and let it cool for five or ten minutes.

Then hold the mouth of the test-tube well under water, and carefully take out the cork. Water rises inside the tube to take the place of the air used up. Notice that the rise is practically one-fifth the volume of air enclosed.

**123. White powder produced when phosphorus burns.**—Carefully dry a wide glass cylinder and a small crucible. Cut off (under water) a piece of phosphorus about half as big as a pea, and dry it between blotting paper. Using a pair of tongs, place the phosphorus in the crucible, touch it with a hot wire, and quickly put the cylinder over it, as in Fig. 79. A white powder is deposited upon the sides of the cylinder. When the phosphorus has ceased to burn, lift up the cylinder and pour a little water into it. The white powder dissolves with a slight hissing noise. Add a little blue litmus to the water: the colour is changed to red.

**124. Use of litmus paper.**—Dip a blue litmus paper into water. No distinct change of colour is observed.

Dip a blue litmus paper into vinegar, lemon-juice, and weak solutions of other acids, such as hydrochloric and sulphuric acids. Notice that each liquid turns a blue litmus paper red.

Dip a reddened litmus paper into a weak solution of soda. Notice that it has just the opposite effect to the acids, and turns the reddened litmus blue again.

**125. Lime-water test.**—Burn a little phosphorus in a test-tube, as in Expt. 122. When the test-tube is cool, take out the cork, and shake up the residual gas with a little clear lime-water which has been poured into it. No change in the lime-water will be observed.

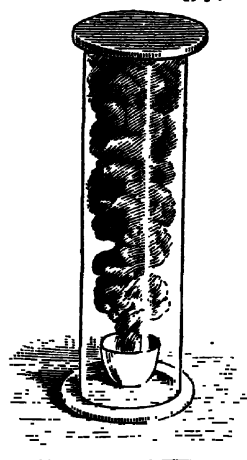


FIG. 79.—Arrangement for collecting the white powder produced by burning phosphorus.

**WARNING! CARE IS NECESSARY IN DEALING WITH PHOSPHORUS. READ THE NEXT PARAGRAPH BEFORE USING PHOSPHORUS.**

**Different kinds of phosphorus.**—There are two kinds of phosphorus, one called yellow phosphorus, and the other red, or amorphous, phosphorus. Yellow phosphorus catches fire

very easily ; the warmth of the hand is quite enough to inflame it. For these reasons it is always kept under water. It is generally bought in the form of sticks, which, when freshly manufactured, are of a yellow waxy colour. This phosphorus can easily be cut with a knife, but the cutting should always be done under water. However small the piece, it must never be touched with the bare fingers, but always lifted by small tongs or forceps. If this precaution is not taken, the warmth of the fingers may cause the phosphorus to catch fire, and, as it is difficult to shake it off when once alight, the burn which it causes is very severe and dreadfully painful. In all experiments with yellow phosphorus these warnings must be borne in mind. The red or powder form of phosphorus is not so inflammable as the yellow kind, and has not to be kept under water, but it must be used with care.

**Phosphorus readily burns in the air.**—It is only necessary to touch a piece of dry phosphorus with a hot wire to make it catch fire and burn. It burns with a dazzling bright flame, and at the same time dense clouds of white fumes are formed, which spread throughout the room. These facts are noticed until all the phosphorus has disappeared.

What happens when phosphorus burns in this way? Is the change anything like that when iron rusts? Does the phosphorus gain or lose in weight? These and several other questions present themselves, and they can be answered now.

**Change produced in air by burning phosphorus.**—To decide whether phosphorus in burning causes the same change in air as iron does when it rusts, it is best to burn some phosphorus in an enclosed amount of air in a way similar to that which has already been described for an experiment with damp iron. One way to do this is to place a little phosphorus on a cork or basin which floats on the surface of water, under a bell jar, or a stoppered bottle having no bottom. After the experiment is over, and the fumes have disappeared, the water is seen to have risen in the jar, indicating that there is less gas in the jar than before the phosphorus was burnt in it.

From what has been previously said, you can understand at once that phosphorus in burning takes out the active part of the air, and leaves the inactive part behind. So far, then, the changes which occur when phosphorus burns are very like

those when iron rusts. Some differences will be studied a little later.

The fraction of the air which disappears as a result of the burning of the phosphorus in a stoppered jar, can be measured easily enough after the jar has been raised a little, so that its mouth is still under water, but it no longer rests on the bottom of the basin. As in the case of the rusting of iron, one-fifth of the air is taken out of it by the phosphorus in burning.

That the gas left behind is really the inactive part of air can be proved by quickly pulling out the stopper of a jar in which phosphorus has been burnt, and introducing a lighted taper. The flame is at once extinguished.

**Phosphorus slowly takes out the active part of the air without being lighted.**—It has been seen that iron slowly takes the active part of the air and combines with it to form rust. And this seems to happen without heating the iron.

Will ordinary phosphorus do the same when it is not alight? This question, too, is easily answered by a simple experiment. When a piece of clean phosphorus is exposed to an enclosed quantity of air over water, the rapid changes just described take place slowly. The only difference in the two cases is the rate at which the active part of the air is taken out. Burning phosphorus combines with the active part very quickly; ordinary phosphorus but slowly. Still, given time enough, ordinary phosphorus will remove all the active part of air, and at the end of the experiment it will be found that again one-fifth of the air has disappeared (Fig. 81).

**Properties of the substance formed when phosphorus combines with the active part of the air.**—We have as yet only noticed that the substance which is formed when phosphorus unites or joins with the active part of the



FIG. 80.—Phosphorus uses up one-fifth of the air in the bottle when it burns, and water rises to take the place of the air used.



air is a white snow-like powder, which corresponds to the rust formed when iron is exposed to air. With a little care, the quick disappearance of the white material which forms the fumes of burning phosphorus can be prevented. All that need be done

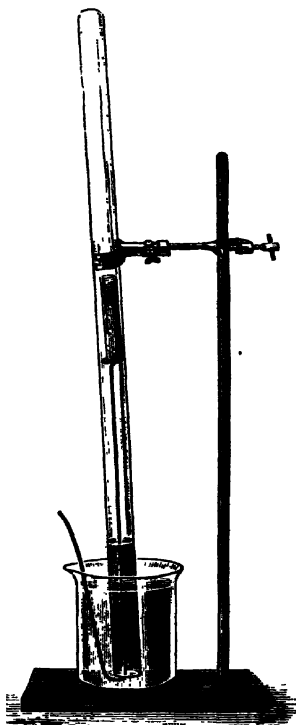


FIG. 81.—If a piece of yellow phosphorus is held at the top of a tube by means of a wire, it will slowly use up the oxygen in the enclosed air, and water will rise in the tube to take its place.

is to burn a piece of dry phosphorus in a dry vessel. In these circumstances, the white fumes settle down on the inside of the vessel in the form of a snow-like solid. But the white powder has so great an attraction for water that so soon as the vessel is opened it extracts the moisture from the air, and, first becoming moist, is quickly replaced by drops of liquid. If water is put into a test-tube or other vessel in which dry phosphorus has been burned, the white powder rapidly dissolves with a hissing noise, like that noticed when water comes in contact with hot iron. When a piece of blue litmus paper is dipped into a solution in water of the white powder produced by burning phosphorus, it is immediately reddened, thus showing that the solution is acid. A solution of soda in water acts in exactly the opposite way, and turns reddened litmus paper blue again.

## BURNING OF A CANDLE

**126. Moisture is formed when a candle burns.**—Over a burning candle hold a clear cold tumbler, which has been carefully dried inside and out. Notice that the inside of the tumbler becomes covered with mist, and, after a short time, drops of water are formed, which run down the sides of the tumbler.

**127. Properties of the gas left after a candle has burnt in air.**—Wind a piece of copper wire round a small candle, as shown in Fig. 83, and light the candle. Push the top of the wire through a small hole in a disc of cardboard, and then lower the candle into a dry, clear glass bottle in such a manner that the top of the jar is covered by the card-

board disc. Observe that the flame of the candle becomes dimmer and dimmer, and soon goes out altogether. Water collects on the inside of the jar, as in the last experiment. Take out the candle, and cover the jar with a greased glass plate. Quickly insert a burning taper, or the re-lighted candle; it is at once put out. Pour in a little fresh clear lime-water, and shake it up in the jar; notice that it is turned milky.

**128. Volume of air used by a candle in burning.**—Fix two or three small candles of different lengths upon the inside of the top of a tin canister. Float the lid upon the surface of water in a basin, or sink it to the bottom of the basin, if the candles project well above the surface while it is in that position. Also float a small dish containing caustic soda. Light the candles, and while they are burning, hold over them a wide-mouthed

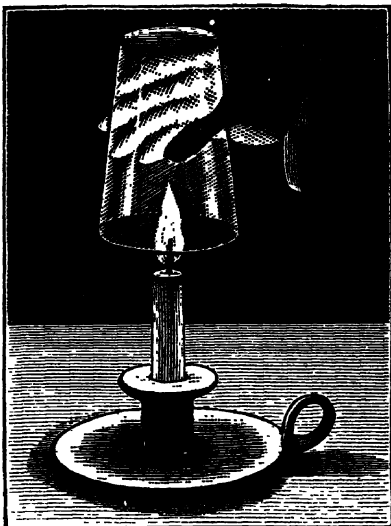


FIG. 82.—To show the formation of moisture by the burning of a candle.

bottle, so that the mouth of the bottle is beneath the surface of the water. When the candles have gone out, wait half an hour, and then mark the place to which the water has risen, by means of a strip of gummed paper. Take out the bottle, and find the volume of water which just fills it. Find also the volume required to fill it to the edge of the gummed paper. Subtract this volume from the preceding one, and thus obtain the volume of air used. Notice that it is *about* one-fifth of the whole volume of air that was in the bottle. One of the products of combustion of the candle is a gas; the caustic soda is used to absorb this product.

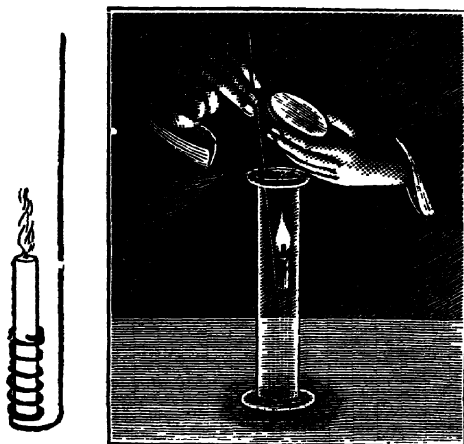


FIG. 83.—Method of supporting a candle for insertion in a bottle.

of the bottle is beneath the surface of the water. When the candles have gone out, wait half an hour, and then mark the place to which the water has risen, by means of a strip of gummed paper. Take out the bottle, and find the volume of water which just fills it. Find also the volume required to fill it to the edge of the gummed paper. Subtract this volume from the preceding one, and thus obtain the volume of air used. Notice that it is *about* one-fifth of the whole volume of air that was in the bottle. One of the products of combustion of the candle is a gas; the caustic soda is used to absorb this product.

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**The burning of a candle.**—Several facts have been learnt about the burning of phosphorus in air, and it will be desirable before proceeding farther to study the burning of some more common combustible substance, such, for instance, as a candle.

In what respects is the burning of a candle similar to the burning of phosphorus, and does it differ from it in any way? It has (p.115) been shown already that a candle will not continue to burn very long in an enclosed quantity of air. Unless the air is renewed in some way, a candle goes out. This gives us a convenient starting-point for the inquiry, Why does the

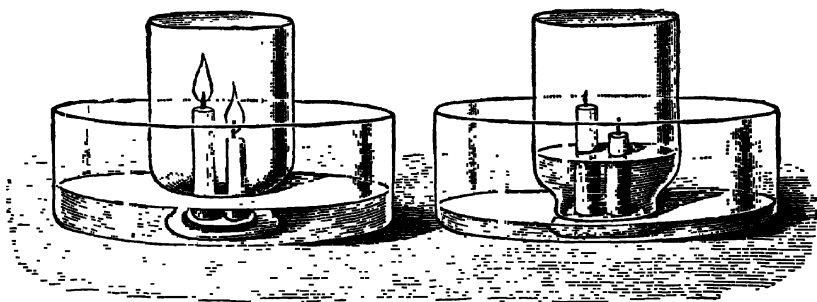


FIG. 84.—When the candles burn, a certain proportion of air is used up, and water rises into the inverted bottle to take its place.

candle go out, and what changes take place when the candle is burning?

**Water is formed when a candle burns.**—When a clear glass bottle, which has been carefully dried inside and out, is held over a burning candle, it is soon noticed that drops of liquid begin to collect on the inside of the bottle and after a time they run down the sides. In some way or other, then, the burning of the candle causes a liquid to be produced. If a sufficient quantity of this liquid is collected, it can be proved to be water by tasting it, or by determining its density, or its boiling and freezing points. Water is the only liquid which boils at  $100^{\circ}\text{C}$ . and freezes at  $0^{\circ}\text{C}$ ., and the density of which is 1.

**Another substance besides water is formed when a candle burns.**—If a candle is burnt in a clear glass bottle in the way shown in Fig. 83, the gas which is left behind can easily be examined. Experiments with this gas show that, like the inactive

part of the air, it will not allow things to burn in it. But besides this it is found that *the gas turns lime-water milky*. The gas left after phosphorus has been burned in a similar jar, has not this property of making lime-water milky; and you can thus be quite sure there is something else in the cylinder, besides the inactive part of the air already described. Hence, when a candle burns, it not only forms water, but also a colourless gas which turns clear lime-water milky. In order that the inactive part of a confined portion of air, left after a candle has burnt in it, may be measured, the gas which turns clear lime-water milky must be removed. It may be absorbed by caustic soda in the manner described in Expt. 128. When this experiment is performed, it is always found that at a certain stage the candle goes out, and subsequent measurement shows that the water has risen and filled about one-fifth of the jar. This result has been noticed in so many lessons that you are thus given reason to believe that the candle is extinguished because it has used up the active part of the air, which makes up one-fifth of the whole.

## BURNING OF COAL-GAS AND LAMP OIL

**129. Moisture and soot produced by flames.**—Invert a clear dry bottle over the flame of a small oil lamp without a chimney. Notice that the inside of the bottle becomes wet. The black deposit which also occurs may obscure the water, but will not hide it entirely.

Hold a flask of cold water over a small gas flame of a laboratory burner (Fig. 85). Notice that the glass becomes slightly dull, owing to the moisture deposited upon it.

**130. Invisible changes caused by burning oil.**—Place a wide-mouthed bottle or a wide glass cylinder completely over a small oil lamp. Notice that after a time the flame goes out. When this happens, quickly remove the cylinder and cover it with a glass plate or a piece of card. Pour in a little clear fresh lime-water. Shake it up, and observe the milkiness produced.

**131. Changes produced by burning gas.**—Bend a glass tube into the shape shown in Fig. 86, and to its long limb attach an india-rubber tube, and so connect the bent glass tube to the gas supply. Turn on the tap,

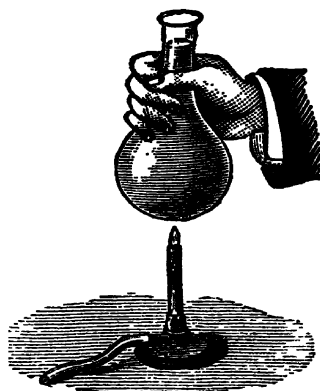


FIG. 85.—How to show that moisture is formed when gas burns.

and light the gas at the end of the short limb. Turn the flame very low, and insert the tube into a cylinder or bottle, and after it begins to get dim, turn off the gas, remove the glass jet from the cylinder, and cover it with a plate. Show that the gas left in the cylinder turns lime-water milky.

**Some familiar combustible bodies.**—Candles are not now commonly used for lighting our houses. Sometimes lamps are employed, and in large rooms ordinary coal-gas is usually the substance which is burnt. As every one knows, lamps are supplied with oil. This oil rises up the wick of the lamp in the

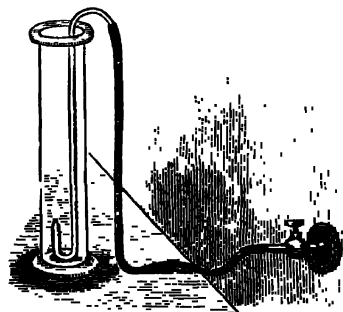


FIG. 86.—How to burn gas in a jar.

same way as water does up cotton threads, and this has been studied already. At the end of the wick the oil is burnt. Are the changes produced by burning oil and gas similar to those which occur when a candle is lighted, and if not, what are the differences?

**What products are formed when oil or gas is burnt?**—By holding a clear dry bottle over the flame of a burning candle, it is easy to prove that water is formed as the

burning is continued, because the vapour becomes condensed on the inside of the cold glass. In exactly the same way, it is found that water is formed when the flame is due to the burning of either oil or gas. This, then, is one way in which these three kinds of burning resemble one another. As in the burning of a candle, so when oil and gas burn, water is formed.

But when a candle is burnt, a gas which turns lime-water milky is formed as well as water. Is this gas also produced when oil and gas are burnt? This question may be answered in the same manner as when studying the burning of the candle. An oil lamp, or a gas jet, or a splinter of wood, is allowed to burn for a few minutes in a glass jar, and is then removed and the jar covered with a glass plate. When lime-water is poured into the jar, it is turned milky. We may consequently say that when a candle, a lamp, coal-gas, and wood are burnt, two substances are formed, namely, water and an invisible gas which turns lime-water milky.

**All substances in burning use up the active part of the air.**—By suitable means it can be shown that oil and coal-gas also unite with the active part of the air and leave the inactive part behind. In fact, this is true in all cases of burning in the air. It does not matter what the substance which burns is like, if it burns in air it does so because it takes out the active part to unite with it to form new substances, and the inactive part is always left behind.

#### SUMMARY.

**Action of metals on air.**—Metals increase in weight when they are tarnished by being heated in the air. Magnesium increases in weight when it burns in air. When iron rusts it increases in weight; the increase is due to a gas taken out of the air by the iron as it rusts. The part of the air left behind as iron rusts will not allow a taper to burn in it.

**Composition of the air.**—Air contains two gases; one is used up by iron in rusting, the other is left behind, and will not allow things to burn in it.

**Proportion of the two gases in air.**—When an excess of iron rusts in an enclosed amount of air, it always takes up one-fifth of the volume of the air. This part, which is also concerned in burning, may be called the *active part* of air. The remaining four-fifths of the air, which will not allow things to burn in it, in which iron will not rust, may be called the *inactive part* of air.

**The burning of phosphorus in air.**—Phosphorus readily burns in the air; in doing so it takes out the active part and combines with it to form a snow-like powder. It can also slowly take out the active part of air without being lighted. When phosphorus is burnt in an enclosed volume of air, one-fifth of this air is used up, and four-fifths remain. The white powder referred to dissolves in water with a hissing sound, forming with the water a solution having *acid* properties.

**The burning of a candle.**—When a candle burns in air, water is formed, and also a gas which turns lime-water milky. As in other cases of burning, one-fifth of the air is used up by a candle in burning. Nothing is lost when a candle is burnt.

**The burning of coal-gas.**—The burning of "gas" is similar to that of a candle. The same substances are produced, and are recognised in the same way. It, too, uses up the active part of air.

#### EXERCISES ON CHAPTER IX.

1. If the inside of a bottle had moist iron filings spread over it, and the bottle were corked up and left in a warm room for a few weeks, what would you expect to see if the cork were then withdrawn while the mouth of the bottle were held under water?

2. Explain the changes that would take place in the bottle mentioned in the previous question.

3. What do you know about the composition of the air? Describe an experiment by which you would show each of the statements you make to be true.

4. Half an ounce of magnesium is burnt in the air, and the ashes formed are carefully collected and weighed. Compare the weight of the ashes with that of the magnesium burnt, and explain the reason for the difference in the weights.

5. Describe an experiment (making a drawing of the apparatus you would use) to show that air contains a gas in which a taper will not burn.

6. A piece of phosphorus is burnt in a stoppered bottle containing 100 c.c. of air. The stopper of the bottle is then taken out under water. How many cubic cms. of water enter the bottle? What facts do you learn from this observation?

7. Compare and contrast the rusting of iron and the slow burning of phosphorus.

8. How could you obtain a small quantity of the white powder formed when phosphorus burns in air? What experiments should you make to show another person the properties of this powder?

9. A piece of phosphorus is placed in a bottle, and the bottle is corked, suspended from one arm of a balance, and counterpoised. The phosphorus is then caused to burn by heating the bottle. Will the balance remain counterpoised after cooling? Give reasons for your answer.

10. A lighted candle is stood upright in water, and a bottle is placed over it, with the neck under the surface of the water. Describe and account for what happens.

11. If two bottles were given you, in one of which phosphorus had been burnt, and in the other a candle, how could you decide which bottle was used for the phosphorus?

12. Write a little essay on "the burning of a candle."

13. Give the names of four common combustible substances. Explain as fully as possible what takes place when a gas jet is lighted.

## CHAPTER X

### COMPOSITION OF THE AIR

#### SEARCH FOR THE ACTIVE PART OF AIR

**132. Changes produced by heating lead in air.**—Heat a few pieces of clean lead in an open crucible (Fig. 87). When the lead has melted, stir the liquid metal with a stout iron wire. Notice the formation of a powdery scum upon the lead. Observe that the colour of the powder is darker when hot. Let the crucible cool. Notice that it now contains a yellow powdery substance in addition to the unchanged lead. By strongly heating this powder its colour changes again, and it becomes red lead.

**133. Gas produced by heating red lead.**—Place a little red lead in a hard glass tube, and strongly heat the tube as in Fig. 88. Notice that the red lead undergoes a change of colour. Into the tube insert a glowing splinter. Observe that the splinter is re-kindled. Why is this?

**134. Change produced by heating mercury rust.**—Repeat the preceding experiment with some red oxide of mercury, and notice the formation of the silvery, mirror-like deposit of mercury, or quicksilver, round the cold upper part of the tube. Insert a glowing splinter of wood, and watch it re-kindle.

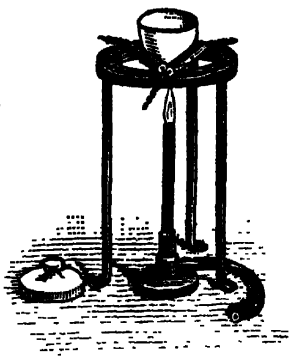


FIG. 87.—Heating lead in contact with air.

**Where to look for the active part of air.**—Since, in suitable circumstances, iron, copper, and lead take the active part out of the air, and unite with it to form fresh substances, it should not be difficult to make these, or similar substances, give up the part of the air which they take up, and so to procure the active constituent of air by itself in a pure form. But a



little more thought suggests that probably some of these substances would do better than others. \*It is quite certain that some are formed more easily than others. Will those which are most easily formed be the best from which to get the active part? No. The reason is this. When a chemical change takes place

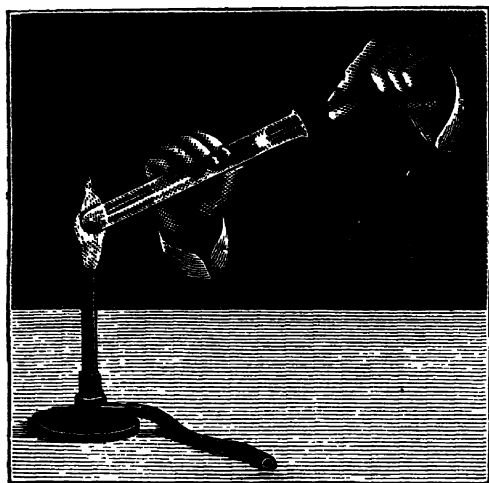


FIG. 28.—When red lead is heated, a gas is given off which will re-ignite a glowing splinter.

very easily, it generally means that the substances taking part in the change have—as it were—a great liking for one another, and when they combine together they form a compound which it is very difficult to separate into its parts again. The easiest way to set to work is, therefore, first to find some substance which only combines with the active part slowly and with difficulty, for the compound such a substance forms with the

active part will most likely be a weak one, and easily broken up again.

**The compounds which lead forms with the active part of the air.**—When lead is heated in contact with the air, a yellow powder which is much darker in colour when hot is formed. If the heating is continued long enough, all the metal is changed into powder. The change takes place fairly easily, so that from previous reasoning you would conclude that it is probably difficult to get the active part of the air again from this powder. And this is so. But it is found that, when some of the yellow powder is heated for a long time at the temperature at which lead melts, it slowly takes up still more of the active part of the air, and changes in colour, becoming red. The first powder obtained, which is yellow, is in some states called *litharge*; the second red powder is known as *red lead*. It is easy to get the second lot of the active part of the air again from the red lead.

**How the active part of air is obtained from red lead.**—When red lead is heated, it changes in colour, and if the heat has not been great, it regains its original red colour when allowed to cool. But if strongly heated, the red lead gives up some of the active part of the air which it contains, and is reconverted into litharge. The amount of the active part of the air which it thus gives up on being heated is the second quantity referred to in the last paragraph, which is taken up slowly when the heating of lead is continued for a long time. If red lead is strongly heated in a tube, as in Fig. 88, and a glowing splinter of wood is pushed down the tube, the splinter bursts into flame and continues to burn brightly. The active part of air has been obtained alone, and supports burning very strongly.

**Other ways of obtaining the active part of air.**—Quicksilver, or mercury, when strongly heated in the air, slowly combines with the active part, and gradually becomes converted into a bright red powder, which is the rust of mercury. If some of this rust of mercury is heated in a hard glass tube, as in Fig. 88, it soon changes in colour; and as the heating is continued it is noticed that a mirror-like deposit is formed round the top, cold part of the tube. When this deposit is rubbed with a penholder or pencil, it runs together and forms little drops of quicksilver. Moreover, if a glowing splinter of wood is introduced into the tube, it bursts into flame, showing that the active part of the air is being driven out of the red mercury rust. This change is just the reverse of what takes place when mercury itself is heated. The active part of the air, which hot mercury slowly combines with, is driven out of the red mercury rust when that is strongly heated. But red mercury rust is very expensive, and it is too costly a plan to heat it to obtain a quantity of the active part of air sufficient for the study of its properties. Several other and cheaper substances easily give up the active part of air when heated.

**The active part of the air is called oxygen.**—As it will be more convenient in the future to speak of the active part of the air by the name chemists use for it, we may state here that it is always called *oxygen*, but the meaning of this name will be better understood later on.

## PREPARATION OF OXYGEN

**135. Oxygen from potassium chlorate.**—Place a little potassium chlorate or chlorate of potash (which is the same thing) in a test-tube, and heat it as in Fig. 88. Observe that the powder crackles, melts, and gives off a gas. Test by a glowing splinter of wood, and see that the gas behaves like oxygen, the active part of the air.

**136. Preparation of a small quantity of oxygen.**—Powder some crystals of potassium chlorate, and mix the powder with a little manganese dioxide (sometimes called pyrolusite). Heat some of the mixture in a test-tube, as in the last experiment. Observe by putting in a glowing splinter that oxygen is given off. Notice that in this case there is no melting, and the gas comes off more readily.

**137. Preparation and collection of oxygen.**—Into a hard glass tube, closed at one end, fit an india-rubber stopper, with one hole in it, through which a tube, bent as in Fig. 89, is passed. The other end of this tube, called the delivery tube, dips under water in a trough. Mix together some potassium chlorate and manganese dioxide, as in the previous experiments, and place the mixture in the tube. Support the tube and delivery tube as shown in the illustration. Fill several bottles with water, and invert them in the trough. Gently warm the tube, and place one of the bottles of water over the end of the delivery tube. As the oxygen is driven off, it displaces the water and gradually fills the bottle. When the bottle is full of oxygen, cover its mouth with a greased glass plate, and lift it out of the trough. In this way fill five or six bottles with oxygen.

**Caution.**—*Be careful not to take away the burner from under the hard glass tube before removing the delivery tube from the trough.*

**138. Physical character of oxygen.**—Take one of the bottles, and examine the gas inside so far as you can by looking at it. Then lift off the glass plate and smell the gas. Draw a little into the mouth, and notice that it has no taste.

**Other substances besides rusts of metals give up oxygen when heated.**—It has been found by experiment that a number of substances give off oxygen when heated. The one most commonly used is a white crystalline solid called potassium chlorate. If a crystal of this substance is placed in a test-tube and heated in the flame of a laboratory burner, it first crackles, then melts, and by and by begins to give off bubbles of oxygen gas. It is easy to show that the gas given off is really oxygen, because when a glowing splinter of wood is inserted into the tube it is immediately re-kindled, bursting into flame and burning brightly.

**Oxygen is given off more easily if the potassium chlorate is mixed with manganese dioxide.**—To obtain

all the oxygen from potassium chlorate requires a considerable amount of heat when the substance is heated alone. If, however, the chlorate of potash is first mixed with certain other substances, such as manganese dioxide, it is found possible to obtain the oxygen with very much less trouble. The gas is given off at a much lower temperature, and more readily in every way. It is the custom, therefore, when it is desired to prepare considerable quantities of oxygen in the laboratory, to use a mixture of potassium chlorate and manganese dioxide, which is often known as *oxygen mixture*.

**Preparation and collection of oxygen.**—The apparatus shown in Fig. 89 is very convenient for preparing oxygen

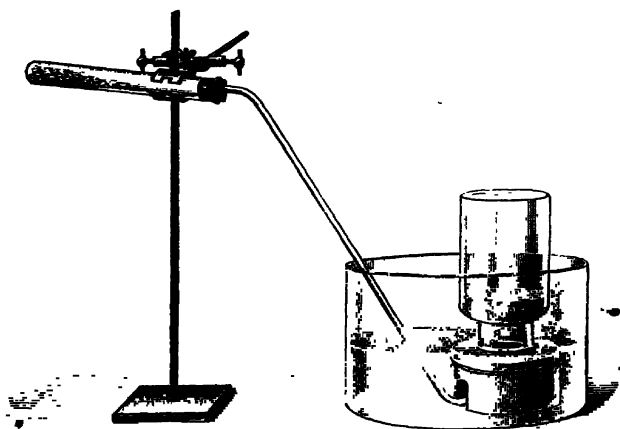


FIG. 89.—A convenient means of obtaining oxygen gas from a mixture of potassium chlorate and manganese dioxide.

from oxygen mixture. A tube closed at one end is fitted with an india-rubber stopper or a good sound cork, through which a hole has been bored. A small glass tube, conveniently bent, fits tightly into the hole, and is so arranged that its other end dips under water in a trough. Oxygen mixture is placed in the closed tube. When this mixture is heated, a gas is given off. To collect this gas, several bottles should be filled with water and inserted in the trough. When these bottles are ready, the closed tube is gradually warmed with a laboratory burner, and after a little while one of the bottles is placed over the open end

of the tube in the trough. The gas which passes down this delivery tube, being lighter than the water, of course rises up through it into the bottle. In this way the water is gradually pushed out of the bottle, and its place taken by the gas.

This plan of collecting gases which do not dissolve, or dissolve only to a small extent, in water, was devised by Priestley, the chemist who first obtained oxygen. It is called *collecting over water*. When the bottle is full of the gas, and while its mouth is still under water, it is covered with a greased glass plate, which is held tightly to the mouth by the left hand, and the bottle lifted out and placed on the table with the right. The gas is now ready for testing.

**Physical properties of oxygen.**—If one of the bottles of gas, collected as described, is allowed to stand for a minute or two and then examined, several of the characters of oxygen can be made out. To begin with, oxygen is an invisible gas. It has neither smell nor taste. It has been found by making careful weighings of the same volumes of oxygen and air that oxygen is a little heavier than air. When made very cold and compressed very much, it is changed into a liquid. All these are called *physical* properties of oxygen.

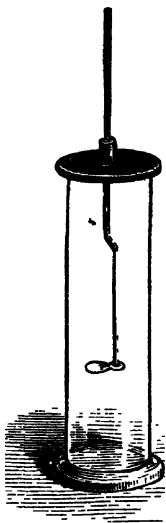


FIG. 90.—A deflagrating spoon in a gas jar.

## CHEMICAL BEHAVIOUR OF OXYGEN

**139. Substances when cold do not burn in oxygen.**—(a.) Place a piece of charcoal in a deflagrating spoon, on the stem of which a disc is placed in such a position that it will cover the mouth of the bottle when inserted in it, while at the same time the spoon is at a convenient distance from the bottom of the bottle (Fig. 90). Quickly plunge this into one of the bottles of oxygen. Nothing happens.

(b.) Make the same test with pieces of sulphur and phosphorus.

**140. Combustion of various substances in oxygen.**—(a.) Repeat each of the previous experiments, using a new bottle of oxygen each time, but first make the carbon red-hot, and ignite the sulphur and phosphorus. Observe that in each case the burning is very much more intense than in air. After the burning has finished, take out the deflagrating spoon and again cover the bottle used with a well-greased glass plate.

(b.) Obtain a piece of iron wire (a thin steel watch-spring will do), and dip one end into a little melted sulphur. Light the sulphur on the wire, and while the sulphur is still burning, place the wire into a jar of oxygen. Notice that the sulphur burns brightly, and then starts the burning of the iron, which continues with a brilliant shower of sparks.

**141. Products of combustion in oxygen.**—*First*, take the bottle in which the charcoal was burnt. Place a piece of moistened blue litmus paper in the bottle. It is turned slightly red. Why? Pour in a little clear lime-water. Notice that it turns milky.

*Secondly*, take the bottle in which the sulphur or brimstone was burnt. Notice the suffocating smell of the fumes. Shake up a little water in the bottle. The fumes dissolve. Put a piece of blue litmus paper into the water. It is turned a bright red.

*Thirdly*, take the bottle in which the phosphorus was burnt. The dense white fumes soon disappear. Test the solution of the fumes with a blue litmus paper. The paper is immediately turned bright red.

*Fourthly*, take the bottle in which the iron was burnt. Notice the brown deposit on the sides of the bottle. Compare it with the rust of iron. Examine the lump of black material found on the bottom of the bottle, or still attached to the remainder of the watch-spring. Is it iron?

**Substances at the temperature of the air do not ignite in oxygen.**—When pieces of cold charcoal, sulphur, and phosphorus are placed into jars of oxygen gas, nothing happens. These observations show that for burning to take place, even in oxygen, we must assist the action by heating the substance until it reaches a certain degree of temperature. For the same reason, if we wish paper, wood, or coal to burn—that is, to unite with the oxygen of the air—we apply heat to them to cause them to do so. If this were not so, since, as you have learnt, the air contains one-fifth of its volume of oxygen, many substances would combine with this gas when simply exposed to the air. The temperature to which a substance must be heated before it will commence to combine with oxygen, or start burning, is often called the *kindling temperature*, or temperature of ignition.

**When heated many substances readily combine with oxygen.**—But if, before putting a piece of charcoal into oxygen, it is made red-hot in the flame of a laboratory burner, or a spirit-lamp, it no sooner comes into contact with the oxygen than a change is noticed. The piece of charcoal begins to burn very brilliantly, bright sparks are given off, and the appearance is so brilliant that the burning puts the observer in mind of a firework. As the burning proceeds, the piece of charcoal

gets smaller and smaller. The charcoal seems to disappear and be lost, but you know that this cannot be so.

Similarly with sulphur, even when a piece is held in a flame, it only catches fire with difficulty, and soon goes out again if removed from the source of heat. But if, while it is still alight, it is transferred to a jar of oxygen, a difference is at once observed. The flame gets larger and brighter, and is of a beautiful lavender colour. It remains until all the sulphur, or all the oxygen, or both, have been used up.

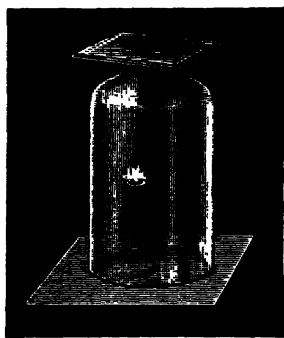


FIG. 91.—Sulphur burns more brilliantly in oxygen than in air.

The flame with which phosphorus burns in oxygen is dazzling in its brightness. In fact, the burning is so intense that it is painful to some eyes to watch it. This is particularly the case when the experiment is performed in a dark room.

**Substances formed when charcoal, sulphur, and phosphorus burn in oxygen.**—*Carbon.*—First, what happens in a jar in which charcoal has been burnt? It has already been stated that the piece of charcoal gets smaller and smaller. An examination of the gas left in the bottle shows that it is not oxygen. It turns a moistened blue litmus paper red; and makes clear lime-water shaken up with it milky. If this solution is placed on one side, a white sediment is formed, and the solution becomes clear again. Evidently, then, when charcoal, which the chemists call *carbon*, burns in oxygen, it gradually diminishes in amount, forming with the oxygen another gas which reddens a blue litmus paper and turns lime-water milky.

*Sulphur.*—Sulphur, or brimstone, disappears as it burns in oxygen. When, however, the contents of a bottle in which the burning has occurred are examined, it is seen that the results are quite different from those in the case of the carbon. The fumes which fill the bottle are very distressing when breathed, their smell is like that produced by a brimstone match when it is first lighted, or by burning sulphur. This is because the gas formed when sulphur or brimstone burns in the air is the

same as when it burns in oxygen. A moistened blue litmus paper put into the gas is immediately turned red, the shade of colour being much brighter and deeper than that produced by the gas in a jar in which carbon has been burnt. Substances which turn a blue litmus paper red are said to be *acid*, consequently the substances formed by the burning of carbon and sulphur in oxygen both have acid properties.

*Phosphorus*.—You are already familiar with the white powder which is formed when phosphorus burns in air, and if it is compared with that produced when phosphorus burns in oxygen, the two are found to be exactly similar. By treating the white powder in the latter case in the same way as that formed when phosphorus burns in air (p. 111), it is found that they are both alike, for each dissolves in water with a hissing noise, and forms a solution which turns blue litmus paper red.

**Burning of iron in oxygen.**—Iron will not burn in the air even when made red-hot. If it did, the iron of fire-grates and furnaces would catch fire and burn away. But iron when made red-hot in oxygen does catch fire, and burns so long as there is any oxygen and iron left. This shows how very much better oxygen supports combustion than air does, which is, of course, what you would expect, since the oxygen in the air is mixed with four times its volume of the inactive part of air—a gas which will not allow things to burn in it at all.

Oxygen is spoken of as a *strong or vigorous supporter of combustion*.

## THE INACTIVE PART OF AIR

**142. The inactive part of air.**—Repeat the experiment of allowing iron to rust in an enclosed quantity of air over water (p. 105), and satisfy yourself that the gas left behind (*a*) extinguishes a flame; (*b*) has no action on a litmus paper; (*c*) does not turn lime-water milky.

**143. Air can be obtained again by mixing oxygen with the inactive part of air.**—Again allow iron to rust in an enclosed amount of air. When the volume of gas in the bottle has ceased to diminish, remove the muslin bag containing the iron. Place the delivery tube from an oxygen apparatus, similar to that used in the experiment on p. 124, and by heating the hard glass tube bubble oxygen into the bottle until it is again full of gas. Cover the mouth of the bottle with a piece of cardboard, lift out the bottle, and test the gas it contains with a burning taper: it behaves just like air.



**The inactive part of air.**—The gas which is left in a bottle of air after iron has rusted in it, or phosphorus has burnt in it, or in which any one of the instances of burning which have now been studied has occurred, will no longer allow a candle or taper to burn in it. This is one reason why it is called the inactive part of air. The inactive part of air does not affect damp iron at all; that is, the iron does not rust when put into it. The name by which this gas is known to chemists is *nitrogen*, and the effect of its presence in the air is that the activity of the oxygen is thereby diminished, toned down, or weakened. Nitrogen not only will not allow things to burn in it, but it will not itself burn, or, as is more commonly said, it is *incombustible*. There are other substances in the air besides the active and inactive constituents which have now been described, but in this chapter, since their amounts are small, their study would be out of place.

**Properties of Nitrogen.**—This gas nitrogen is an example of a very inactive, or inert, substance. It can only be made to combine with other substances with great difficulty—in fact, with most other things it cannot be made to combine at all. Not only is nitrogen incombustible, and a *non-supporter of combustion*, but it will not turn lime-water milky, and a mouse put into the gas dies very quickly—this is because of the absence of oxygen.

If the number of negative properties of nitrogen is borne in mind, and if we think of this side by side with the very active qualities of oxygen, it is not at all difficult to understand that its presence in the air serves the purpose of weakening the oxygen, and thus making combustion less intense.

**Chemical composition of air.**—What may be termed the fundamental gases in air are oxygen and nitrogen. Carbon dioxide and water vapour are practically always present, and various other gases, or vapours, frequently occur in small quantities, but these may be regarded as impurities, and ought not to be considered as constituents of pure air. The following table shows the percentage composition of air as regards volume; that is, for instance, the number of cubic feet of the various gases present in 100 cubic feet of the atmosphere:—

<i>Oxygen</i> , a gas which supports combustion . . . . .	20·61
<i>Nitrogen</i> , an inert gas . . . . .	77·32
<i>Carbon dioxide</i> , a suffocating gas . . . . .	0·03
<i>Water vapour</i> . . . . .	1·40
Argon . . . . .	0·63
Nitric acid } . . . . .	traces
Ammonia } . . . . .	
Ozone } . . . . .	

These proportions are remarkably constant in ordinary air, and it is only in localities or under conditions of an exceptional character that they vary to any noteworthy degree. In the air of mines, the *oxygen* has been found as low as 18·6 per cent., but this represents almost the lowest percentage of oxygen ever obtained from a place where human beings could live. In the midst of vegetation, or open ground, especially in the daytime, oxygen is present in the proportion of nearly 21 per cent., but never more.

*Nitrogen* not only serves to tone down the activity of oxygen as a supporter of combustion, but it is also useful to the life of plants. A few lowly plants appear to absorb nitrogen direct from the air, but the majority of them obtain it indirectly as the result of the action of bacteria existing in the soil or in their roots.

The proportion of *carbon dioxide* rarely exceeds 3 parts in 10,000 in pure air, and is not often less than 2·7 parts per 10,000. During the night the proportion is slightly greater than in the day. In the streets of a town the amount of carbon dioxide only exceeds the average amount of the open country by about 1 part in 10,000. In rooms, however, and badly-ventilated places, carbon dioxide is often greatly in excess, and oxygen is present in a much smaller proportion than it ought to be. Carbon dioxide is not essentially a poisonous gas, but it is often found in bad company, and when it occurs in excess, the air of which it forms a part is unfit to breathe.

Air always contains a certain proportion of invisible *water vapour*, and when the air is cooled to a sufficient degree, this vapour becomes visible in the form of mist, fog, cloud, rain, or other familiar forms of water (p. 99). *Ozone* is a peculiar form of oxygen, and is usually present in the air of the open country or over the sea, but not in that of towns. Traces of

ammonia are always found in air, and with nitric acid gas it plays a part in the growth of plants.

In addition to these and other gases, numerous minute solid particles are suspended in air, some of them being living germs. They are more abundant in the town than in the country.

### SUMMARY.

**The active part of air.**—To obtain the active part of air by itself those substances are again heated which metals form when heated in the air. Such substances as are formed with difficulty are split up most easily. *Red lead* is a good instance. The *red oxide of mercury* is also easily split up into mercury and the active part of air. Chemists call the active part of the air *oxygen*.

**Preparation of oxygen.**—When *potassium chlorate* is heated, it crackles, melts, and gives off bubbles of oxygen. If the potassium chlorate is first mixed with *manganese dioxide*, the oxygen is given off more readily. This mixture is often called *oxygen mixture*.

**Physical properties of oxygen.**—It is a gas with *no colour, no taste, and no smell*. It is a little *heavier than air*.

**Chemical behaviour of oxygen.**—At the *temperature of the air*, carbon, sulphur, phosphorus, &c., *do not burn* when placed in oxygen. If these substances are *first heated*, they burn very brightly if plunged into oxygen. In each case the substance formed has *acid* properties. Iron will burn in oxygen under suitable conditions.

**The inactive part of air.**—The gas left behind when iron rusts in air is called *nitrogen* by the chemist. It extinguishes a flame, has no action on litmus paper, and does not turn lime-water milky. A mixture behaving just like air can be made by mixing nitrogen and oxygen in the proper proportions. Nitrogen serves to *dilute* the oxygen of the air.

**Composition of the air.**—In 100 parts by volume of air, the following gases are usually present, in the proportion stated:—

Oxygen, 20·61 ; nitrogen, 77·32 ; carbon dioxide, 0·03 ; water vapour, 1·40 ; argon, 0·63 ; and traces of nitric acid, ammonia, and ozone.

### EXERCISES ON CHAPTER X.

1. Why can oxygen be more easily obtained by heating red lead than iron rust?

2. How can oxygen be obtained from red oxide of mercury? Draw the apparatus you would use to perform this experiment.

3. How would you set about obtaining two or three bottlesful of oxygen? Make a sketch of the apparatus you would employ, and state any precautions you would observe.

4. If two bottles were given you, one containing oxygen and the other ordinary air, how could you find out which was which?

5. What happens when a piece of cold charcoal is put into oxygen?

What differences would you notice if the charcoal were first made red-hot, and then plunged into the oxygen?

6. How would you show by an experiment that iron can be made to burn in oxygen? •

7. If you were given a bottle in which phosphorus had been burnt, and another in which carbon had been burnt, how could you distinguish one from the other?

8. Compare the properties of nitrogen and oxygen. Tell what you can about the use of each of these gases in the atmosphere.

9. What do you mean by the physical properties of a gas? State the physical properties of nitrogen.

10. How would you obtain a bottle full of nitrogen? What experiments should you perform to demonstrate its chief properties?

11. Mention some everyday chemical changes in which the oxygen of the air takes a prominent part.

12. What useful purpose is served by painting iron fences?

## CHAPTER XI

### WATER AND ITS CONSTITUENTS

#### ACTION OF METALS ON WATER

**144. Iron becomes rusty in water containing dissolved air.**—Leave a piece of clean iron in water for several hours. Notice that rust forms upon it.

**145. Iron in boiled water does not rust.**—Boil some water in a flask or clean saucepan for some time so as to drive out the dissolved air. Obtain

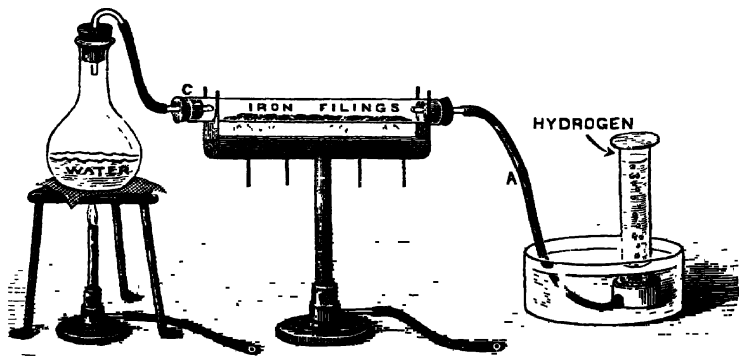


FIG. 92.—A means of obtaining hydrogen gas from water by depriving steam of oxygen.

a wire nail, or a steel pen-nib, and after washing it in soap and water, to remove oil, dry it, place it in a bottle, and quickly fill up the bottle with the boiled water. Insert a tightly-fitting india-rubber stopper, being careful to enclose no air between the stopper and the top of the water. Place on one side, and examine the iron after several hours. No rust will be seen.

**146. Inflammable gas obtained from water.**—Place some clean iron filings or iron tacks in a hard glass tube connected at one end with

a flask containing water, which has boiled for some time, while the other end is connected with a tube leading to a trough of water (Fig. 92).

Strongly heat the iron filings and boil the water in the flask, so that steam passes over the heated iron, and then into the water in the trough, where it condenses. Now place over the end of the tube in the trough an inverted test-tube or gas jar of water. Notice the steam seems not completely condensed, but that small bubbles ascend to the top of the gas jar. In this way collect a jar full of the gas. When the gas jar is full of gas, pull out the stopper at one end of the tube, and remove the burner from underneath the flask of water. While the gas jar still dips under the water in the trough, cover its open end with a greased glass plate, and, holding a lighted match to the mouth with your left hand remove the glass plate. Observe that the gas takes fire with a slight pop, and *burns*.

Now examine the iron filings in the tube, and observe that a quantity of rust has been formed.

**Action of cold iron upon water.**—Whenever iron is left in water for some time, a considerable quantity of reddish rust is formed. It has, however, already been learnt that water contains air dissolved in it, and also that iron can combine with the oxygen in the air. This suggests, therefore, that iron which rusts when left in water, does so by combining with the oxygen of the air which the water has previously dissolved.

When water is boiled for some time, the air which it has dissolved is driven out. If a piece of bright clean iron is sealed up in a tube, with some of this boiled water from which all the air has been driven, and left for some time, it is found either that the iron does not rust at all, or that it does so to a very small extent only. Evidently, then, little can be found out about the composition of water by putting cold iron into it. But when heated iron is used, a different result is obtained.

**Action of heated iron upon water in the form of steam.**—To make as large a surface of heated iron as possible come into contact with the water, it is best to use iron filings in studying the action of steam upon hot iron. The filings are contained in a hard glass tube, which will not melt when strongly heated. The hard glass tube is fitted at each end with india-rubber stoppers, through each of which a hole is made to take a glass tube. One of these tubes is connected with a flask, or other vessel, containing boiling water. The other tube is bent to make a delivery tube, which dips into a trough of water.

The iron filings in the hard glass tube are made red-hot by strongly heating the tube. The tube is then connected with the flask containing boiling water, and the stopper with the delivery tube is inserted at its other end. At first, air escapes through the water in the trough; and, after that, bubbles of steam pass from the end of the delivery tube and are condensed by the water in the trough. When this occurs, a tube full of water is inverted over the end of the delivery tube, and as it is watched it is noticed that small bubbles of gas pass up into the tube and collect there. When a lighted match is brought in contact with the gas which collects in the test-tube, it is found to burn. Moreover, an examination of the iron filings shows they have been partly changed into red rust.

**What happens when steam passes over red-hot iron.**—Since the gas which collects in the test-tube burns, it is clearly not uncondensed steam, for when a lighted match is put into steam the steam does not burn. Nor is it air from the flask, or from the water, because air will not burn. It must be a new substance which has been obtained from the water, and you can now be quite sure that the water contains a gas which will burn. This gas used to be called *inflammable air*, that is, air which will burn. It is now called *hydrogen*.

But why should the iron have rusted? What has already been learnt about the rusting of iron in air must be borne in mind. Iron, when it rusts in air, takes out the active part, that is the oxygen, and combines with it. There is every justification for saying, therefore, that the iron filings have combined with oxygen, or they would not have gone rusty. Where did the oxygen come from? Certainly not from air, because that was driven out before the change took place. The result indicates that the oxygen comes from the steam, and that steam is made up of this oxygen and the gas which burns, that is, hydrogen. It is, however, necessary to study the question more fully before it is established that this conclusion is correct.

## THE COMPOSITION OF WATER

**Caution!** *Sodium must never be touched with bare fingers, for it becomes very hot when in contact with wet or moist things. Experiments with sodium should only be performed in the presence of some one familiar with its properties.*

**147. Effect of putting sodium in water.**—Cut off a piece of sodium about the size of a pea. By means of forceps, place it upon the surface of some water in a glass tumbler, and quickly put a large glass beaker or a wire-gauze dish-cover over the tumbler. Notice that the sodium floats, takes the form of a ball, and runs about over the surface, and finally bursts with a slight explosion.

Bring a lighted taper near a piece of floating sodium. A golden yellow flame appears, and the sodium appears to be burning.

**148. Sodium used to set free the inflammable gas in water.**—Make a short roll of wire-gauze, close one end, and insert a pebble or a small piece of lead into the roll or tube. Cut off a piece of dry sodium about the size of a pea. Put it into the roll, and then nearly close the open end. Have an inverted test-tube full of water ready in a trough or a basin, and drop the gauze, with the sodium in it, into the water. Notice the bubbles of gas which rise from the sodium, and collect them in the test-tube. Examine the gas in the test-tube, as described in Expt. 146, and notice that it burns.

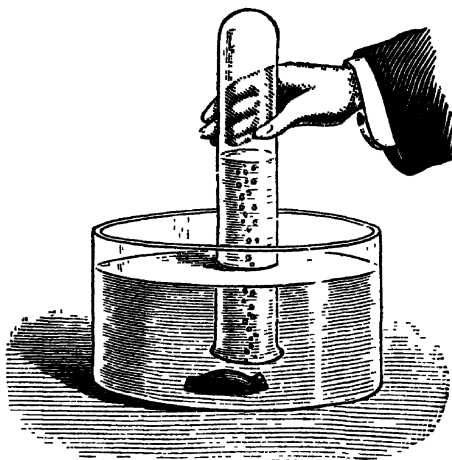


FIG. 93.—Hydrogen gas driven out of water by sodium.

**Other metals besides iron will turn hydrogen out of water.**—It has been seen that before iron can set hydrogen free from water it must be red-hot. Some other metals will do the same thing without being heated. It has been found that a metal called *sodium* will turn hydrogen out of water even when it is cold.

**The action of sodium on water.**—Though sodium is heavier than naphtha (the liquid in which it is always kept), it is lighter than water. Consequently, when a piece of sodium is put upon water it floats. So soon as it touches the water, however irregular its shape, the piece of sodium becomes quite round, and darts about from one place to another, spinning the whole time. All the while it gets smaller and smaller, and just before its disappearance there is often



a little explosion, and small pieces of sodium are scattered in every direction. The greatest care must be taken that none of these tiny fragments hit the face.

If, as the little ball of sodium is careering about the surface of the water, a lighted taper is brought near it, it seems to catch fire and burn with a golden yellow flame. It is, however, not really the sodium which is burning, but the inflammable gas which it sets free from the water as it travels from place to place.

**How to collect the gas which sodium turns out of water.**—If a piece of sodium is put into a weighted roll of wire-gauze, which is dropped into water, then, so soon as the sodium enters the water, bubbles of gas begin to appear. These can easily be collected by holding an inverted test-tube full of water over the roll of gauze. In this way tubes can be filled with the gas which sodium turns out of the water.

**The gas which sodium turns out of water will burn.**—When one of the tubes which has been filled with the gas which sodium turns out of the water is covered with the thumb of the left hand before it is lifted out of the water, and while still held mouth downward a lighted taper, held in the right hand, is brought near to its mouth, on removing the thumb the gas burns with an almost colourless flame. It seems to be the same sort of gas as was got from steam by the action of heated iron: it really is so.

## PREPARATION OF HYDROGEN

**149. Production of inflammable gas by the action of zinc upon an acid.**—Select a flask, or wide-mouthed bottle, and fit it up as is shown in Fig. 94. Be very careful that both the stopper and the tubes fit very closely. Into the flask put enough granulated zinc to cover the bottom. Pour some water upon the zinc. Arrange the delivery tube in the trough as shown.

Pour a little sulphuric acid down the thistle-headed acid funnel, and be quite sure that the end of the funnel dips beneath the liquid in the flask. Do not collect bottles of the gas until you are sure all the air has been expelled, which you can find out in the following way:—

Fill a test-tube with water, and invert it over the end of the delivery tube. When it is full of gas, still holding it upside down, take it to a flame, which should not be near the flask you are using. Notice that there is a slight explosion. Continue this until the hydrogen burns quietly down the test-tube.

When this happens, you may proceed to fill one or two bottles. When the bottles have been filled, it is better not to remove them from the water until you want to use them.

Collect also a soda-water bottle half full of the gas.

**Caution.**—*Be careful not to bring a light near the thistle-funnel, or tube delivering the gas, even when the action in the flask seems to have ceased, or a dangerous explosion may occur.*

*Be careful also that none of the acid used gets upon your fingers or clothing.*

**Usual method of preparing hydrogen.**—The gas which collects when steam is passed over red-hot iron comes off slowly, and it takes a long time to collect a sufficient amount of it to enable its properties to be easily examined. It would

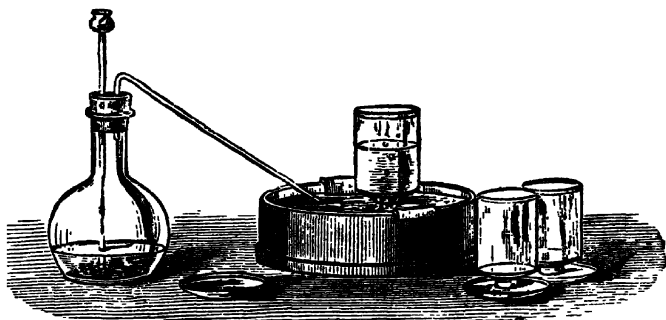


FIG. 94.—When dilute acid is poured upon granulated zinc, a gas is given off and can be collected in bottles.

take too long to obtain bottles full of hydrogen by acting upon water with small pieces of sodium. When larger quantities of hydrogen are wanted quickly, dilute sulphuric acid is poured upon pieces of zinc; the zinc gradually dissolves in the acid; at the same time bubbles of a gas which burns, and is hydrogen, are given off very rapidly. This method of making the gas is generally made use of when enough hydrogen is required to examine its chief properties.

A flask, or wide-mouthed bottle, is fitted up as shown in Fig. 94. The india-rubber stopper or cork with two holes through it must fit very closely. The thistle-headed acid funnel and the delivery tube, too, must pass tightly through the holes in the cork, for hydrogen at once finds out any leak,

however small. Enough granulated zinc to cover the bottom of the bottle is put into it, and is covered with water. The end of the acid funnel must dip below the water when the cork is put into the neck of the bottle, or the hydrogen may escape by its means.

**Collection of hydrogen.**—So soon as the dilute sulphuric acid is poured down the acid funnel and reaches the zinc, hydrogen begins to come off. Before pouring it in, therefore, there must be something ready for collecting the gas. Some bottles are filled with water and inverted in the trough of water. A small amount of dilute sulphuric acid is then poured down the funnel, and the gas at once starts coming off, but *none of it must be collected until all the air has been expelled and the hydrogen given off is quite pure.* You have seen that this can be found out by filling a test-tube, over the water in the trough, with the gas in the usual way. To begin with, a slight explosion occurs when the test-tube of gas is held near a flame. But after several test-tubesful have been collected and tested in the same way, it will be found that the gas burns quietly down the tube. When this happens, one or two bottles may be filled with the hydrogen in the same way.

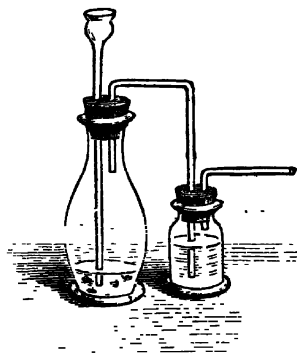


FIG. 95.—How to purify hydrogen before collecting it.

## PROPERTIES OF HYDROGEN

### 150. Simple examination of hydrogen.

—Observe as many properties of hydrogen as you can by examining one of the bottles full of the gas. Notice that it is colourless and has no taste.

Pure hydrogen has no smell, but the gas when prepared from ordinary sulphuric acid and zinc has a slightly disagreeable smell. This is due to the presence of small quantities of another gas. This impurity may, however, be removed by passing the gas, before collecting it as described in the last section, through a bottle arranged as in Fig. 95, containing a little Condyl's fluid.

(The jars, or bottles, when taken from the trough, must be held mouth downwards as shown in Fig. 94.)

**151. Inflammability of hydrogen.**—Test one jar of the gas by means of a lighted match or taper, as in Fig. 96. Observe that the gas burns around the mouth of the jar, that the flame of the taper is extinguished,

but when taken out it again becomes alight on passing through the flame of the burning hydrogen.

**152. Lightness of hydrogen.**—Take a jar or cylinder, A (Fig. 97), containing air only, and hold it mouth downwards in one hand. Take

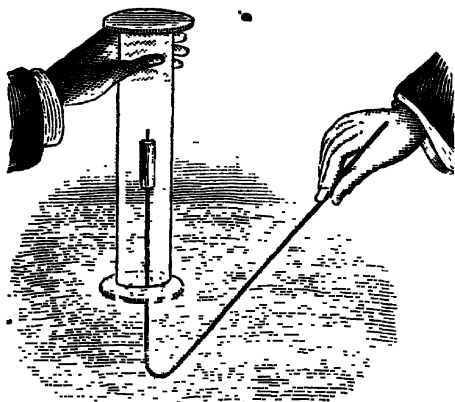


FIG. 96.—The hydrogen burns at the mouth of the bottle, but the taper is extinguished when inside the bottle of hydrogen.

another jar or cylinder, filled with hydrogen, in the other hand, and bring it mouth downwards by the side of the first, then gradually bring it into position B. Now put the lower cylinder down, and apply a lighted taper to each jar. Show, by applying a light, that the jar which contained air now contains hydrogen, which passed upwards from B.

**153. Explosion of a mixture of hydrogen and air.**—

Take a strong soda-water bottle (egg-shaped), and half fill it with hydrogen, as described on p. 139. Wrap the bottle well in a duster, and hold it firmly. Take the

bottle out of the water, so that the water in it runs out. The bottle is now filled with a mixture of air and hydrogen. Apply a light, and you will not fail to observe that an explosion results.

**154. The flame of burning hydrogen.**—

Fit a straight tube drawn out to a point to a hydrogen generator, as shown in Fig. 98. After pouring a little sulphuric acid down the thistle-funnel, collect a test-tube of the gas issuing from the straight tube, and hold the mouth of the test-tube near a flame, *which must be a few feet away from the generator*. The gas will at first go off with a pop, or burn with a squeaking noise, but after two or three trials it will burn quietly, with a blue flame. *When you can carry this flame of burning hydrogen to the apparatus from which the gas is being produced, do so, and use it to ignite the gas escaping from the pointed tube.* If you

remember always to do this, there can be no danger, for when you are able to carry a flame of hydrogen in a test-tube of the gas, for a distance of two or three feet to the generating apparatus, you may be sure that the hydrogen issuing from the apparatus is not mixed with air.

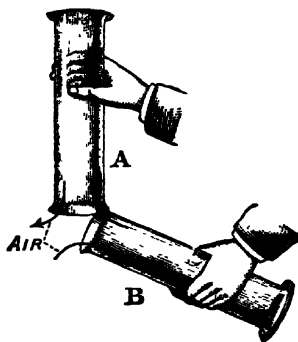


FIG. 97.—Pouring a gas upwards. The hydrogen from B passes up into A, and the air from A is driven out.

Light the hydrogen in this way.

Observe that it burns with a pale blue flame, which after a time becomes yellow. This coloration is due to the glass becoming hot, and some of the substances in it being burnt in the flame of hydrogen.

**Physical properties of hydrogen.**—Hydrogen is a gas which possesses no colour, no taste, and no smell. It is very much lighter than air. For this reason soap-bubbles, or balloons filled with it, rise in the air when free to move. It has been found by many carefully performed experiments that hydrogen is the lightest substance known. The weight of a unit volume, or litre, of hydrogen is consequently taken as the standard with which chemists compare the densities of other gases.

**Hydrogen will burn, but it will not allow things to burn in it.**—When a lighted match is brought to the end of a delivery tube, from which pure hydrogen is issuing, the gas catches fire and continues to burn with a flame which is at first almost colourless, but very soon changes to a bright golden yellow colour, which is due to the sodium contained in the glass. Hydrogen is therefore said to be *combustible*.

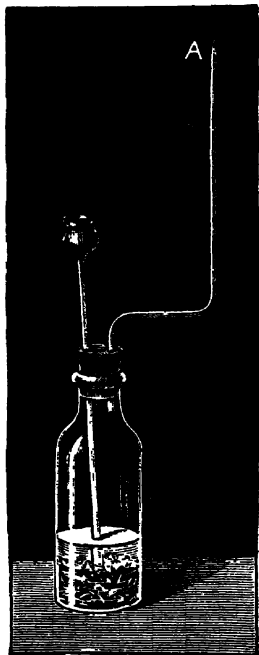


FIG. 98. — Arrangement for obtaining a flame of hydrogen.

If a jar of hydrogen is held with its mouth downwards, and a lighted match (or a piece of taper) fixed to the end of a wire is pushed up into the jar, the hydrogen, as in other cases, catches fire, and continues to burn at the mouth of the jar; but the flame of the match or taper is put out, showing that things will not burn when surrounded by hydrogen gas. If the match or taper is slowly drawn out of the jar, it will be again ignited as it passes through the hydrogen flame at the mouth of the jar. Thus, though hydrogen is a combustible gas, it *will not support combustion* in ordinary circumstances.

**Hydrogen being lighter than air can be poured upwards.**—When you wish to pour water from one vessel to another, you know very well that the vessel *into* which water is to be poured is placed below, and the vessel *from* which the water is to be poured is placed above and tilted. This is because water is heavier, bulk for bulk, than air. Similarly, if it were required to pour a gas considerably heavier than air from one vessel to another, like the gas which is produced when carbon is burnt in oxygen, you would proceed in just the same way as with water. Since, however, hydrogen is much lighter than air, the order of things must be exactly reversed. The vessel *into* which hydrogen is to be poured must be placed above and be inverted, while the vessel *from* which the gas is to be poured must be below and held with its mouth upwards. This method of pouring hydrogen from one jar to another is shown in Fig. 97, and is always used when it is required to pour a gas lighter than air from one vessel to another.

## COMPOUND FORMED BY BURNING HYDROGEN

**155. Liquid formed by burning hydrogen.**—Arrange a flask or bottle, as before, for the production of hydrogen. Pass the gas through a tube containing chloride of calcium, in order to dry it thoroughly. Fill a test-tube with the gas, as already described, and when the gas will burn, use the flame to ignite the gas issuing from the tube, and allow it to burn under a retort, which is kept cool by a stream of water flowing in where the stopper is generally put, and out at the end of the neck. Observe the formation, on the outside of the retort, of a clear liquid, which collects and drops into a clean beaker placed to receive it (Fig. 99).

In this way obtain sufficient of the liquid to enable its properties to be examined. Much time is saved if a number of pupils all work at the same experiment, and the liquid formed by all the hydrogen flames is added together.

**156. Nature of liquid formed by burning hydrogen.**—As already learnt, proceed to find the boiling point, freezing point, and density of the liquid formed when hydrogen burns in the air. The boiling point is found to be  $100^{\circ}\text{C.}$ , the freezing point  $0^{\circ}\text{C.}$ , and the density 1. The liquid is consequently pure water.

Evaporate some of the liquid in a basin, and notice that no residue is left.

Test some of the liquid with litmus paper. Observe that it is neutral.

**The burning of hydrogen.**—When an apparatus is arranged as in Fig. 99, so that a hydrogen flame burns near a

surface which is kept cold, it is found that a liquid is continuously formed, which, as the experiment proceeds, collects drop by drop in a beaker put to catch it. Since this liquid is formed by the burning of a perfectly dry gas, it must be a substance which hydrogen forms with something out of the air. To be sure that dry hydrogen is burnt, the gas is first passed through a tube containing lumps of calcium chloride, a substance which has the power of absorbing moisture.

**The liquid formed as hydrogen burns is water.**—When the liquid produced by burning dry hydrogen in air is examined, it is found to have all the properties of pure water.

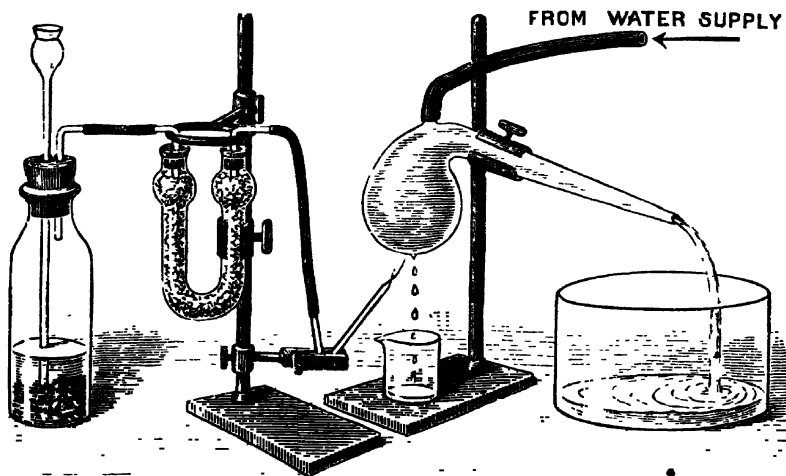


FIG. 99.—Method of obtaining and collecting water by burning hydrogen gas in air.

Thus, its boiling point is  $100^{\circ}$  C., its freezing point  $0^{\circ}$  C., and its density is 1; when evaporated to dryness, no residue is left behind, and it does not change the colour of a litmus paper used for testing it.

These, as you know, are all characteristics of pure water, and it is therefore quite certain that the liquid which is formed as hydrogen burns is water.

**Hydrogen in burning uses up one-fifth of the air.**—When hydrogen burns in an enclosed quantity of air over water, it is found that after a time the flame goes out, and the water

risers in the vessel in which the air was enclosed to take the place of that part of the air with which the hydrogen combines. If the amount of water which entered the flask in this way is measured, it tells the amount of oxygen in the air in the flask. By filling the flask with water and measuring it, the volume of air the whole flask holds can be found. On comparing the amount of water which rises into the vessel with the capacity of the vessel, it is found to be one-fifth.

This proves that hydrogen, like other combustible substances, combines with the active part of the air when it burns, and you see that the conclusion arrived at in the last paragraph was quite correct.

**Hydrogen and oxygen will not combine at ordinary temperatures.**—If no light is near, hydrogen gas may escape into the air without any danger, and with no accompanying chemical change. This shows that the two gases, hydrogen and oxygen, do not combine at ordinary temperatures. If they did so unite, every bubble of hydrogen which came into contact with the air would catch fire, and form a small quantity of water by combining with the oxygen of the air. The same facts are noticed when hydrogen and air are mixed in a soda-water bottle. The gases do not affect one another, but when a flame is brought near to the mixture combination at once takes place and a loud explosion is heard. The heat of the flame raises their temperatures enough to cause the two gases to combine.

## WATER CONTAINS HYDROGEN AND OXYGEN

**157. Action of hydrogen on heated copper oxide.**—Arrange an apparatus like that shown in Fig. 100, in which a flask A for making hydrogen is connected with the bottle B containing strong sulphuric acid. The passage of the hydrogen through the strong acid completely dries the gas. A small amount of the black oxide of copper is placed in the hard glass tube AC which is about 1·5 centimetres in diameter, and fitted with the tube D in the manner shown. When you are sure that pure dry hydrogen is escaping from the open end of D, heat the oxide of copper in the tube by means of the burner E. In a few minutes moisture will be seen to collect in the tube D, and presently to drop into the dish put to collect it.

Examine the residue in the hard glass tube, and note its change to a reddish colour, due to the presence of metallic copper.

Rearrange the apparatus, using the U-tube G in the place of the tube D. The U-tube G, containing lumps of calcium chloride, is employed to collect the water formed.



**Water is formed when hydrogen is passed over heated oxide of copper.**—When pure hydrogen, which has been dried by passing it through some substance which very readily absorbs moisture, such as strong sulphuric acid, or calcium chloride, is passed over heated black oxide of copper, the oxide is found to lose in weight, giving up something to the hydrogen. With this “something” the hydrogen forms water. The oxide of copper is at the same time changed into a reddish powder which on examination is found to be copper.

We cannot be wrong, therefore, in supposing that the some-

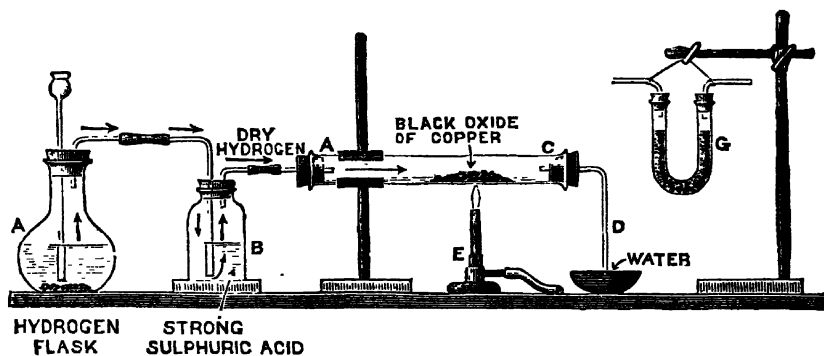


FIG. 100.—When hydrogen is passed over hot oxide of copper, it extracts the oxygen to form water, and leaves copper behind.

thing taken out of the oxide is the active constituent of the air, with which copper unites when it changes into copper rust or oxide. Another name for this active constituent is, as you know, oxygen. Hence, it may be said that hydrogen in passing over heated copper oxide takes oxygen from it, and combines with it to form water. Or, water is composed of hydrogen and oxygen.

#### SUMMARY.

**Action between water and iron** —If bright iron is left in water it becomes rusty, but if all the air is first driven out of the water, the iron does not rust. When steam is passed over heated iron, the iron rusts, and an inflammable gas known as hydrogen can be collected. Iron rust has been seen to contain oxygen, and it seems that iron takes oxygen out of the steam to form rust, and that the hydrogen collected

is another constituent of the steam. Steam apparently consists of hydrogen and oxygen.

**Action of sodium on water.**—When sodium is thrown on water, it takes oxygen out of the water, combining with it to form a rust of sodium, and drives out of the water an inflammable gas, *hydrogen*. The rust of sodium dissolves in the water, forming a solution which turns a red litmus paper blue.

**Larger quantities of hydrogen** are prepared by acting upon sulphuric acid with zinc. It must always be borne in mind that a *mixture of hydrogen and air is highly explosive*.

**Properties of hydrogen.**—When pure, hydrogen has no colour, no taste, and no smell. It is the *lightest* gas known. It burns in the air; but burning substances put into it are extinguished.

When dry hydrogen is burnt in air, it unites with oxygen to form water.

Hydrogen and oxygen will not unite at ordinary temperatures.

When hydrogen, which has been freed from moisture, is passed over heated copper oxide, it unites with the oxygen of the oxide to form water.

### EXERCISES ON CHAPTER XI.

1. Describe how, if you were provided with some iron tacks, you could obtain hydrogen from steam.

2. What can be learnt as to the composition of water by passing steam over heated iron?

3. How would you show that the gas obtained by passing steam over heated iron is the same as that turned out of water when a piece of sodium is thrown upon it?

4. State your reasons for believing that water contains hydrogen and oxygen.

5. Give reasons why, in preparing hydrogen by means of zinc and sulphuric acid: (a) the bottom of the thistle funnel must be under the liquid in the flask; (b) the apparatus must be air-tight.

6. How would you proceed to obtain a few jars full of pure hydrogen?

7. Describe some experiments which show that hydrogen is lighter than air.

8. How would you show: (a) that hydrogen will burn? (b) that hydrogen will not allow substances to burn in it?

9. In what respects is the burning of a jet of hydrogen like, or unlike, the burning of a candle?

10. What substances are formed when hydrogen burns in air? How could you collect and test the products of combustion?

11. First, some copper is heated to redness in a crucible freely exposed to the air; what happens? The solid is next taken out of the crucible, placed in a hard glass tube, heated, and dry hydrogen passed over it. What exactly happens in this case?

## CHAPTER XII

### NATURAL WATERS

#### CONSTITUENTS OF NATURAL WATERS

**158. Evaporation of different natural waters.**—Measure out 250 c.c. (that is  $\frac{1}{4}$  litre) of clear rain-water, tap-water, and sea-water, into evaporating basins, the weights of which have been determined. Evaporate to dryness in each case on a sand-bath, taking care to lose no

liquid by spurting. When the evaporation is complete, and the residue (if any) left behind is dry, allow the basins to cool, and again weigh. The increase in weight of each basin represents the amount of dissolved material in equal volumes of each of the kinds of natural water.

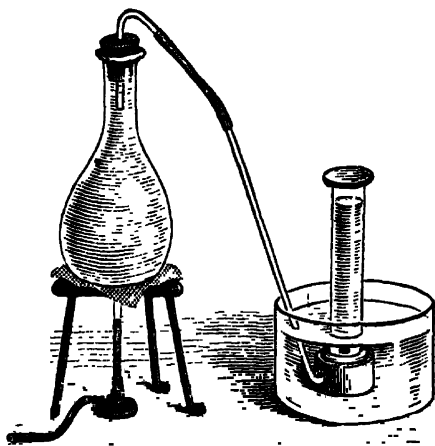


FIG. 101.—Gases driven out of water by heat.

**159. Purification of water by distillation.**—Distil some tap-water and sea-water, as described on p. 85. Observe that in both cases pure, or distilled water, is obtained.

**160. Gases dissolved in water.**—Completely fill a gallon tin can or a large flask with tap-water, and attach a cork and delivery tube which has also been filled with water, dipping the tube under an inverted cylinder of water over a pneumatic trough (Fig. 101). Heat the flask by means of the flame of a laboratory burner, and observe that the dissolved gases are driven out of the water. Collect these gases in the ordinary way. When you have enough, test, first, with a little lime-water, which being turned milky shows there is carbon dioxide in the gas; and, then with a lighted taper. Since the flame is not extinguished, the gas collected also contains oxygen.

**161. Filtering through sand.**—Clean a flower-pot, and plug the hole at the bottom with a piece of sponge. Place sand, with pebbles or

charcoal, in the pot, as shown in Fig. 102. Support the pot on a tripod stand, and place a tumbler beneath the hole. Pour on to the top some muddy water from the gutter, and notice that the water which trickles through and is caught in the tumbler is comparatively clear.

**Natural waters.**—Pure water does not occur naturally on the earth. This is because of the great solvent power which water possesses. At the moment rain is formed it consists of pure water, being really distilled water which is formed in just the same way as when this liquid is prepared in the laboratory by the condensation of steam. The heat of the sun converts the water from lakes, rivers, seas, and other bodies of water, into vapour, which in the higher or colder parts of the atmosphere is cooled and again forms water. Yet no sooner is the rain formed than it begins to dissolve various substances. In its passage through the air it takes up varying amounts of the gases which make up the atmosphere, such as oxygen and the gas called carbon dioxide, which is formed whenever carbon combines with oxygen, either in burning or in the body. When the surface of the earth is reached, the water dissolves portions of all the soluble ingredients out of the soil and the underlying rocks. The most soluble bodies are, of course, dissolved to the greatest extent. The solvent power of water is considerably increased by the presence of the carbon dioxide it dissolves, partly from the air and partly from the soil.

**Spring or tap water.**—Spring and tap water are often the same liquid. Tap-water is simply that which is supplied to our houses from the large iron pipe or water main which is buried in the road. The water from the reservoir, which the main contains, is sometimes only spring-water which has been brought from a distance. Its *dissolved* substances have not been

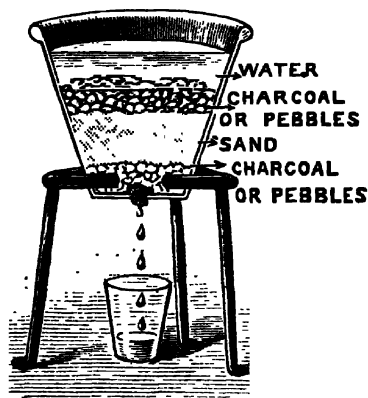


FIG. 102.—A simple means of filtering water.

removed from it, though its *suspended* impurities have generally been got rid of by filtering. This filtering, which has to be done on a very large scale, often takes place through beds of sand.

When the properties of, and dissolved substances in, such tap-water are examined, therefore, a knowledge of spring-water is obtained. The amount of residue left in the basin when tap-water is evaporated depends upon where the water was brought

from. Some tap-water, such as that which has been drawn from a chalk country, like parts of Kent, Sussex, and Hampshire, contains a great deal of dissolved solids; while that of Manchester, for example, which comes from the Lake district, where the rocks are hard and only dissolved with difficulty, leaves very little residue when evaporated. Though tap-water which contains a small amount only of dissolved material is the best for washing purposes, the dissolved substances in other specimens of tap-water are not necessarily injurious to health when used to drink. On the contrary they are often beneficial; it is said that water

which contains no chalk dissolved in it is not good for children, because they require this substance in considerable quantities to build up their bones.

**Sea-water.**—As everybody knows, sea-water is often called *salt water*. The water of the ocean contains sufficient dissolved matter in it to make it taste, while in spring, rain, and river water the quantity is so small that the water is tasteless; for this reason the water of springs and rivers is often called *fresh water*.

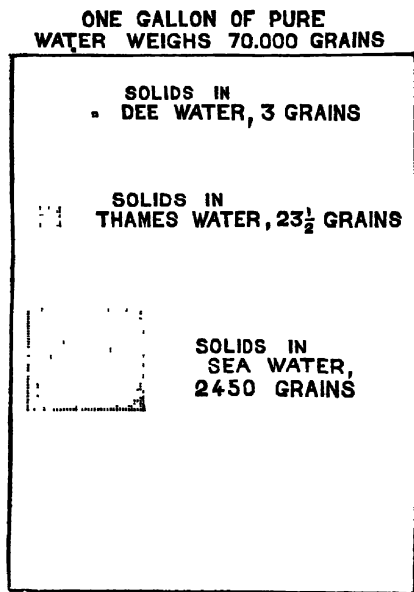


FIG. 103.—Comparative weights of a gallon of pure water, and of the solids found in river and sea waters.

Streams and rivers are continually flowing into the sea, and carrying into it the substances they have dissolved out of the land. Evaporation is going on unceasingly from the surface of the ocean, robbing it of pure water in the form of water vapour, with the result that there is an accumulation of soluble material in the water which is left. It would consequently seem at first that the sea must every day become saltier. This is certainly not the case. It would be so if it were not for the animals and plants which live in the ocean, and extract from the water some of its dissolved material to build up their solid parts.

**Gases dissolved in water.**—When water is heated in a glass vessel, minute bubbles are seen on the inside of the glass. These are bubbles of gas driven out of the water. By suitable means the gases dissolved in water can be collected and analysed, and when this is done it is found that they consist chiefly of oxygen, nitrogen, and carbon dioxide, which are the essential constituents of air, though the proportions in which they occur are different. Air does not, in fact, dissolve in water as a whole, but its constituents are dissolved separately, each according to its solubility. A litre of river water contains about 4 to 8 cubic centimetres of oxygen, 12 to 18 cubic centimetres of nitrogen, and 2 to 25 cubic centimetres of carbon dioxide.

The oxygen dissolved in river-water is necessary for the life of fishes and other aquatic animals. When a river receives a large amount of refuse of various kinds, oxygen is used up during the decay of this matter, and is frequently diminished to such an extent that fishes die in the water from suffocation. The water of the Thames shows very clearly the difference in the proportion of oxygen at various points. Above London, it contains about 7 cubic centimetres of oxygen per litre; at Hammersmith the proportion at the same time has been found to be less than 5 cubic centimetres per litre; at Somerset House—close to London City—only 1·5 cubic centimetres of oxygen per litre occurs, and at Woolwich only 0·25, that is,  $\frac{1}{4}$ , of a cubic centimetre per litre. Several miles below London, after more gas has been absorbed, the proportion of oxygen again rises, and fishes are able to live in the water.

**Impurities in water.**—Oxygen, nitrogen, and carbon dioxide are not impurities in water; in fact, water which does not contain these gases dissolved in them is flat and insipid.

Mention has already been made of the mineral substances dissolved in water, but even these, though often undesirable, do not always pollute water. There is evidence that water of excessive hardness may interfere with the healthy working of certain parts of the body in the course of time, while, as has been mentioned, water with little or no chalk dissolved in it is said to be bad for children ; but, with these exceptions, hard and soft waters, if free from deleterious organic substances, are equally wholesome.

The sources of danger in water are the organic substances present. They are of animal or vegetable origin. They are not very easily recognised by a simple chemical test, and water which is suspected to be bad can only be tested satisfactorily by an analytical chemist.

In addition to the chemical examination, suspected water should also be examined for the presence of the micro-organisms of diseases. Water frequently becomes infected with disease germs from sewage and other sources of contamination, and if it is drunk it may be the means of producing an epidemic. Typhoid fever and cholera are two diseases which are known to be communicated to people by water polluted with the germs of these diseases.

**Purification of water.**—Water supplied to towns is first filtered through sand, which removes from 93 to 99 per cent. of the bacterial life. Upon this subject, Professor Percy Frankland, a distinguished authority upon micro-organisms in water, says :<sup>1</sup> "The river Thames, in its raw condition, before undergoing any treatment at the hands of the water companies, has on the average throughout an entire year, as many as 20,000 micro-organisms in one cubic centimetre (or about 20 drops) of water. Now, it used to be formerly supposed that the processes of purification as carried on at the works of the water companies, which consist in storage in large reservoirs with subsequent filtration through sand, were of little if any use at all. But on examining this same water after it has been in the hands of the water companies, I found that it contained, on delivery from the mains, in the same number of drops, on an average for the same period of time, only 400. We have here, then, the efficiency of sand filtration put practically to the test, and the assurance that

<sup>1</sup> *Our Secret Friends and Foes* (S.P.C.K.).

it forms a very material protective measure against our infection by water-carried microbes."

**Filters.**—Water may become infected with disease germs after its passage from the sand-beds of the water companies, so that there is some justification for filtering it when it is received from the tap. Water which is obtained from wells or springs is so liable to contamination that it ought always to be filtered or boiled before drinking. The ordinary filter is, however, worse than useless so far as removing harmful microbes is concerned. The carbon, spongy iron, or similar material through which the water passes, removes sand or other suspended particles, but it can no

more prevent the passage of disease-germs than a sieve can stop water from passing through it. Moreover, unless the material used in the filter is frequently cleansed or changed, it becomes a veritable forcing-house for micro-organisms and other organic im-

purities, and so the last condition of the water of the filter is worse than the first. The perfect filter is almost a "castle in the air," for if the rate of filtration is rapid, the filter is not effective, and if it is effective as regards the prevention of the passage of bacteria, it is too slow of action to be of much practical use. The best filter is that known as the Pasteur-Chamberland, in which the water filters through

a cylinder of unglazed porcelain. This is made in a variety of shapes, the most useful being made to fit upon the tap of the water service, so that the pressure of the mains forces the water

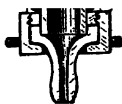
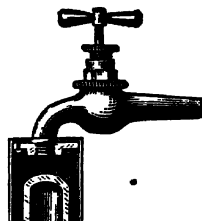


FIG. 104.—Pasteur-Chamberland Filter, attached to a water service tap.

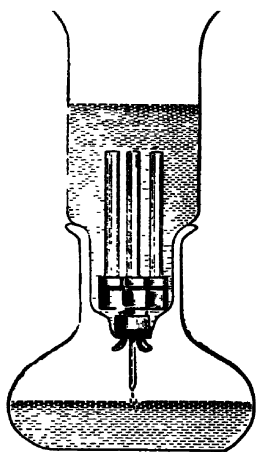


FIG. 105.—Pasteur-Chamberland Table Filter, with decanter to collect the filtered water which passes through the porcelain cylinder.



through it at the rate of about two or three quarts an hour (Fig. 104). When this pressure is not utilised, the water only passes through the filter in drops. For instance, a table filter of the form shown in Fig. 105, and having a capacity of three pints, will filter about four pints of water per day, or a little more, according to the number of times it is refilled.

In the absence of a filter, the best precaution to take in the case of water suspected to be bad is to boil it. Boiling for five minutes is sufficient to destroy practically all micro-organisms ; indeed, most of such organisms are killed by a temperature considerably lower than that at which water boils.

### HARD AND SOFT WATERS

**162. Solution of chalk in water.**—Pass carbon dioxide (obtained from an apparatus like that shown in Fig. 106 containing pieces of

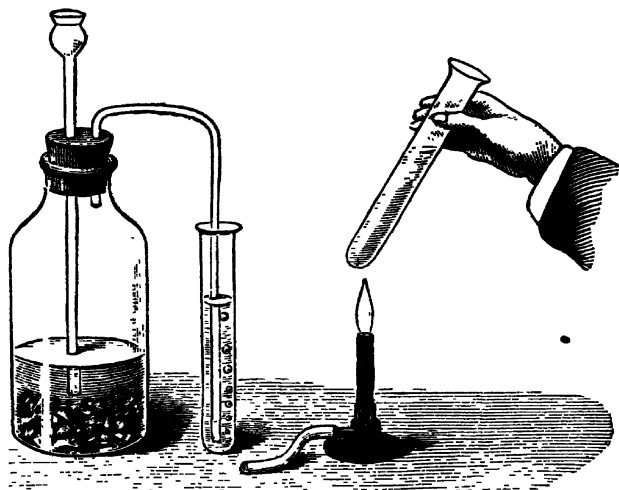


FIG. 106 —When carbon dioxide is passed into clear lime-water, the water turns milky, and afterwards clear. The milkiness is again obtained by boiling the lime-water.

marble on which hydrochloric acid has been poured) into clear lime-water until the solution is clear again, after first becoming milky.

**163. Effect of boiling water in which chalk is dissolved.**—Boil a part of the clear solution in a test-tube. Observe the return of the milkiness. What is the cause?

Boil in a flask a little natural water which is hard because of the presence of chalk; observe that the water becomes milky.

**164. Soap solution to test the hardness of water.**—Dissolve some Castile soap in spirits of wine and label the solution "soap solution." Add a few drops of soap solution to distilled water. Shake the mixture. Notice a lather is easily formed.

**165. Temporary and permanent hardness of water.**—(a) Add soap solution to some of the clear solution obtained by passing carbon dioxide for some time through lime-water (Expt. 162). Shake the mixture, and observe that a lather is formed only after a considerable quantity of soap solution has been added.

(b) Similarly, observe that it is difficult to make a lather with a solution of calcium chloride.

(c) Boil portions of the solution in the last two experiments before adding the soap solution. Soap lathers in the first solution quite easily after it has been boiled; boiling makes no difference to the second solution.

**166. Hardness of sea-water.**—Try whether the soap solution forms a lather when shaken up with sea-water. Does boiling make any difference?

**167. Measurement**

**of the hardness of water with soap solution.**—(a) Pour some soap solution into a burette. From this add soap solution (a drop or two at a time) to about a wine-glassful of rain-water in a bottle; insert the stopper and shake vigorously after each addition. When a lather forms which remains for five minutes, cease to add soap solution, and note the volume of it used.

(b) Treat an equal quantity of spring-water in the same way and note the amount of soap solution required.

(c) By taking various samples of water, equal in quantity, and treating them as described, you may roughly compare the facility with which they form a lather. The *harder* the water the more is the amount of soap solution required.

**168. Hard water softened by the addition of lime.**—Into water



FIG. 107.—Carbon dioxide is produced by the action of dilute acid upon "fur" from a kettle, and may be poured from the flask into a beaker containing lime-water, which it turns milky.

which contains dissolved carbonate of lime pour clear lime-water until the water becomes quite milky. Then filter. Find how much soap solution is required to produce a lather which remains for five minutes in 50 c.c. of the clear water. Test an equal volume of the water which has not been treated with lime-water in the same way.

**169. "Fur" from a kettle compared with chalk.**—(a) Collect some of the crust or "fur" from the inside of a tea-kettle; compare its appearance with that of chalk.

(b) Put a small quantity of clear lime-water into the bottom of a tumbler or test-glass. Put some of the "fur" into a small flask, and pour on it a little dilute hydrochloric acid. Note the effervescence which takes place. Tilt the flask over the tumbler so that the escaping gas will flow from the flask to the tumbler, taking care that no liquid runs out of the flask. Shake the tumbler containing the lime-water. Note that the latter turns milky (Fig. 107).

(c) Put a little of the fur into a wide test-tube and pour upon it dilute acid. Plunge a lighted taper into the tube, or a burning wood-splinter. Note that it is extinguished.

Perform similar experiments with chalk; notice similar results.

**170. How to obtain a deposit of "fur" from spring-water.**—Put a short piece of clean glass tubing into a kettle which is in constant use. At the end of a week or two examine it. Break it, and notice that fur is only deposited on the outside.

**Hard and soft waters.**—It is a fact familiar to every one that soap lathers very easily in some waters and not at all in others. If rain-water be used the lathering takes place with great ease, while with the water supplied to some towns a lather can only be made with difficulty; and if we attempt the same process in sea-water there is no lathering at all. *Those waters in which soap lathers easily are said to be soft. When this is not the case, the water is spoken of as hard.*

**Soft waters.**—Water containing no dissolved mineral matter, or water which is pure, dissolves soap easily, and when the solution is shaken up a lather is at once formed.

**Hard waters.**—Hard waters always contain dissolved mineral substances. These substances may either be soluble in pure water itself, or only in water which has first dissolved carbon dioxide. When water is boiled it is hotter than when it is taken from a spring or tap, and as hot water usually dissolves more of a substance than cold, you would hardly expect mineral substances in water to be given up when the water containing them boils. But this does happen, and the reason is that boiling drives out the carbon dioxide from solution in the water.

Certain solids which are soluble in water owing to the presence

of carbon dioxide, are not dissolved by pure water ; hence when the water is boiled and the gas is driven off, these substances are deposited upon the sides of the vessel in which the boiling is done.

These facts enable us to divide hardness in waters into two classes—(1) Hardness which *is not* removed by boiling, called *permanent*; (2) Hardness which *is* removed by boiling, called *temporary*.

**How natural waters become hard.**—No sooner is the water vapour in the air condensed into rain than its work of solution begins. First, the rain in its passage to the earth dissolves portions of the gases in the air. The most soluble of these gases are oxygen and carbon dioxide, and evidently there will be most of these dissolved. When the surface of the earth is reached, the solution of many more substances begins. The amount of each material dissolved, of course, depends upon its solubility.

Of these substances dissolved some are soluble in pure water, others dissolve only in water containing carbon dioxide. The hardness to which the first class gives rise is *permanent*, that caused by the second class *temporary*.

**Ways of getting rid of temporary hardness.**—First, there is the plan of boiling, which has been already mentioned. But it is quite clear that when engineers wish to soften the supply of water to a town they cannot boil the whole amount. Another process is adopted. The object in view is to get rid of the carbon dioxide in the water. If this can be done, some of the substances in solution in the water become insoluble, and are given up. What the engineer often does is to have lime put into the water. The lime combines with the carbon dioxide to form chalk. The carbon dioxide being removed from the water, it can no longer dissolve the carbonate of lime, and this, together with the chalk formed, is thrown down.

**Ways of getting rid of permanent hardness.**—Boiling is of no use in removing hardness when the water is permanently hard. But there are two other ways of softening permanently hard water.

The first plan is employed by the washerwoman. She puts soda into the water. The chemist's name for soda is sodium carbonate. When the sodium carbonate comes in contact with

the substance which causes permanent hardness, a chemical change takes place, and the originally soluble substance causing the hardness is changed into a compound insoluble in water, which is consequently thrown down. The water is thus softened. The substance in solution in water which usually causes permanent hardness is calcium sulphate, one form of which is plaster of Paris. When washing-soda comes into contact with this compound, the calcium sulphate is changed into calcium carbonate, the chemist's name for chalk, which, being insoluble in pure water, is thrown down as a precipitate, and so removed.

The second plan is called *distillation*, and is equally useful in both kinds of hardness.

**Reason for getting rid of hardness.**—One object in view in getting rid of every kind of hardness is to diminish the amount of soap necessary in using water for cleansing processes. When soft water is used, the soap immediately dissolves in it, causing a lather at once. When hard water is employed, a considerable quantity of soap is first used up in combining with the chalk and other substances, which render the water hard, to form the curdy material seen at the top of hard water after soap has been used in it. This curdy matter is insoluble in water, and represents so much wasted soap.

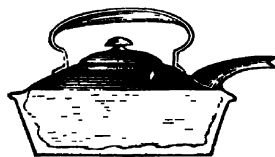


FIG 108.—The deposit of fur in a kettle consists chiefly of chalk.

**Furring of kettles.**—An examination of the inside of a kettle which has been in use for some time generally reveals the presence of a white chalk-like deposit. If a little of this white incrustation is removed and examined, it is found to be brittle, and when treated with

an acid it effervesces and gives off a gas which puts out flames. But this is just what takes place when an acid is poured upon chalk. The fur inside the kettle is chalk. It was present in the water which has been from time to time boiled in the kettle. The solubility of the chalk depended upon the carbon dioxide dissolved in the water, and when the water was boiled the carbon dioxide was driven out, and the water was therefore unable to dissolve the chalk, which was consequently thrown down on the inside of the kettle to make the white deposit.

## SUMMARY.

**Rain-water** has been naturally distilled, and is, therefore, at the moment of its formation, pure water—*free from dissolved matter*.

In its passage through the air it dissolves oxygen, nitrogen, and carbon dioxide. The last-named aids the water in dissolving solid matter from rocks and soils.

If the rocks through which the water passes contain easily soluble substances, the water will contain much dissolved material, while water which passes over hard insoluble rocks contains little solid matter in solution.

**Sea-water** contains a much larger amount of dissolved matter than *spring-water*. While rivers pour their waters into the sea, the sea-water is evaporated, and the dissolved solids remain behind.

Plants and animals are continually removing dissolved solids from sea-water.

**Dissolved solids.**—A gallon of pure water weighs 70,000 grains. A gallon of filtered Dee water contains only 3 grains of dissolved matter in it. A gallon of filtered Thames water contains  $23\frac{1}{2}$  grains, and a gallon of filtered sea-water 2450 grains ( $5\frac{1}{2}$  ozs. nearly).

**Dissolved gases.**—A litre of river water contains about 4 to 8 cubic centimetres of oxygen, 12 to 18 c.c. of nitrogen, and 2 to 25 c.c. of carbon dioxide.

**Impurities in water.**—Dissolved solids and gases usually do not make water unwholesome. Organic substances, whether animal or vegetable, make the water deleterious—especially if such impurities are of the nature of the micro-organisms of disease.

**Filtering.**—Filtering through sand greatly reduces the number of micro-organisms, and is very much to be recommended.

Most filters used in the household are worse than useless. The best is the Pasteur-Chamberland filter. To boil water for five minutes is sufficient to destroy practically all micro-organisms.

**Soft waters** are those in which soap lathers easily.

**Hard waters** are those containing dissolved solids which use up soap without making a lather.

**Hardness in water** is of two kinds: (a) *Permanent*, caused chiefly by the presence of calcium sulphate, which can be got rid of only by adding washing-soda, or by distillation. (b) *Temporary*, caused chiefly by carbonate of lime, and got rid of by boiling or by adding lime.

**Hard water requires more soap** to form a lather than soft water; the mineral substances unite with the first portion of soap to form a useless curd.

**Furring** of kettles and boilers is due to the deposition of carbonate of lime from the water, when carbon dioxide is driven off by boiling.

## EXERCISES ON CHAPTER XII.

1. Spring and river water are not pure. Describe how you prove this and determine the quantity of dissolved solids in a gallon of each.
2. Sea-water has a disagreeable taste. Why is this? How do you

account for the facts that : (a) All the rivers run into the sea, yet the sea is not full. (b) The amount of dissolved solids in sea-water does not increase?

3. How would you proceed to show that : (a) Rain-water is free from dissolved solids. (b) Sea-water contains very much dissolved material?

4. Describe an experiment to show that hard water requires more soap to form a lather than soft water.

5. What is the cause of hardness in water, and how may it be removed?

6. What is the reason for sometimes adding washing-soda to water? How does the soda affect the water?

7. Why are waters sometimes hard, and in what ways does their hardness differ?

8. After washing my hands in spring-water with soap, I noticed a curd or scum on the water; account for this. What would have happened had I used rain-water, and why?

9. River-water always contains dissolved gases. What are these gases? Whence do they come? What is usually the amount present? What useful purpose do all, or any of them, serve?

10. What are the usual impurities of river-water? Distinguish between an impurity and a dissolved constituent of water.

11. What are the advantages of filtering a town's supply of water through beds of sand?

12. Discuss the advantages and disadvantages of ordinary household filters.

13. What procedure would you adopt if you had reason to suspect that the water supply to your house was contaminated with organic impurities?

14. Draw and describe the Pasteur-Chamberland filter.

15. Account for the "fur" which sometimes collects in kettles and boilers.

16. Why are some waters described as "hard" and others as "soft"? Which are better for cleansing purposes, and why?

# CHAPTER XIII

## CARBON DIOXIDE

### BURNING OF CARBON

**171. Carbon in organic substances.**—Heat a series of organic substances, such as meat, wood, potato, egg, &c., in a crucible, and notice in all cases the production of a black residue, consisting largely of carbon. Heat more strongly, and observe that it burns away, leaving an almost colourless ash.

**172. Carbon dioxide produced by burning charcoal and wood.**—(a) Heat strongly a piece of charcoal in a closed hard glass test-tube, and show that without air it does not burn.

(b) Suspend a piece of glowing charcoal in a bottle containing lime-water. Shake up, and show that the lime-water is turned milky. Carbon dioxide can always be distinguished by this action upon lime-water.

**173. Carbon dioxide in air.**—Pour some clear lime water into a blue dinner plate, or some other shallow vessel of a dark colour. Leave it exposed to the atmosphere for a little while. Notice the thin white scum formed on the top. The carbon dioxide in the air has turned the top layer of liquid milky.

**174. Character of air changed by breathing.**—Fit two bottles with corks and tubes, as shown in Fig. 109. See that the corks are air-tight. Put some clear lime-water into each bottle. Place the tube C, or an india-rubber tube leading from it, in your mouth. When you suck at the tube, air is drawn in through the glass tube which dips into the lime-water in the bottle A. When, however, you blow instead of sucking, your breath passes out through the tube which dips into the lime-water in the bottle B. Notice that the lime-water in A remains

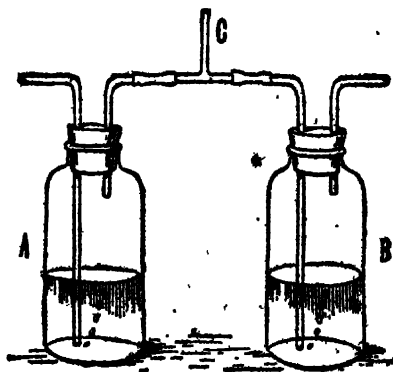


FIG. 109.—Arrangement for showing that carbon dioxide is among the gases breathed out from the lungs.



almost clear, but that in B is rendered milky by the air you breathe out. You thus see that fresh air has little effect upon lime-water, but breathed air quickly turns clear lime-water milky.

**175. Oxygen from plants.**—Take a bunch of fresh watercress, or water-weeds, and put it into a beaker or glass jar completely filled with water saturated with carbon dioxide. Cover the plants with a funnel nearly as wide as the jar, as shown in Fig. 110. Fill a test-tube with water, and invert it over the funnel. If properly managed there should at first be no gas in the test-tube. Place the jar in bright sunlight for an hour or two, and then examine it. You will notice bubbles of a gas have collected at the top of the tube. Test the gas with a glowing splinter of wood. It is found to be oxygen.

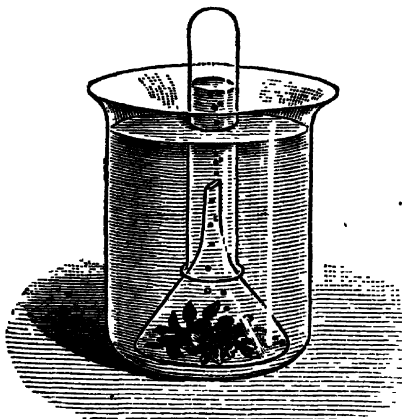


FIG. 110.—Oxygen is given off from fresh green plants in sunlight, and can be collected as shown.

**176. Plants in sunlight and in darkness.**—Repeat the whole experiment, but, instead of putting the bottle in bright sunlight, place it in

the dark. Observe that in such circumstances no bubbles of oxygen are formed.

**Gas formed when any kind of carbon, or any combustible substance containing carbon, is burnt in air.**—When things such as candles, oil, gas, and wood are burnt either in the air or in pure oxygen, a colourless gas with no smell is produced, which has the power of turning lime-water milky. All these substances contain, in one form or another, a substance called carbon. The gas produced when these substances burn is called carbon dioxide, or carbonic acid gas. The same gas is produced by burning carbon in air or oxygen. In fact, whenever a substance rich in carbon burns in a plentiful supply of air or oxygen, this carbon dioxide is produced. Knowing how many fires there are in all civilised countries, in houses, furnaces, engines, and so on, it is not difficult to understand that at every hour of the day very large quantities of this carbon dioxide are formed, which escape, sooner or later, into the air.

Carbon dioxide gas is the only colourless gas with no smell which has this power of turning lime-water milky, so that this fact forms a convenient way of showing the presence of carbon dioxide.

**Carbon dioxide is given off in breathing.**—If a person blows with the mouth into clear lime-water, the lime-water is turned milky (Fig. 111). This is another important fact. It is clear that carbon dioxide escapes from our lungs in breathing. And so it does from every animal. Not only, then, do all cases of ordinary burning result in the addition of carbon dioxide to the air, but also every act of breathing. It does not matter how small the animal is, all the time it is alive it is continually adding to the atmosphere a certain amount of a suffocating gas, the presence of which can be detected by its action in causing turbidity in clear lime-water.



FIG. 111.—Clear lime-water is turned milky when a person breathes into it.

**Chemical combustion in the body.**—Since breathing adds carbon dioxide to the air, surely carbon and oxygen unite together in some way in the body to produce this gas. What really happens is something like this. When air is drawn into the mouth, the mixture of oxygen and nitrogen of which the air is composed passes down the wind-pipe into the lungs inside the chest. The lungs, which are like the “lights” or lungs of animals to be seen at the butcher’s, are two spongy masses which take up a large part of the chest (Fig. 112). They contain throughout a very large number of little air spaces with very thin walls. All over these thin walls tiny blood-vessels are spread, and the walls of these are also very thin. The oxygen of the air passes through these thin walls into the blood, which carries the oxygen to every part of the body. The number of little air-cells in the lungs is estimated at five or six millions, and their surface at about twenty square feet.

Whenever you lift an arm, move a leg, or use any muscle, part of the substance of the muscle is used up or burnt. It

joins with the oxygen in the blood. The muscle contains, with other elements, hydrogen and carbon. The hydrogen of the muscle unites with the oxygen brought by the blood to form water; and the carbon of the muscle combines with some more of this oxygen to form carbon dioxide. These waste products—water and carbon dioxide—are taken away from the muscles by the blood, which is always in motion, to the lungs. Here the carbon dioxide is given up, and passes, with nitrogen and other gases which you breathe out, back into the atmosphere.

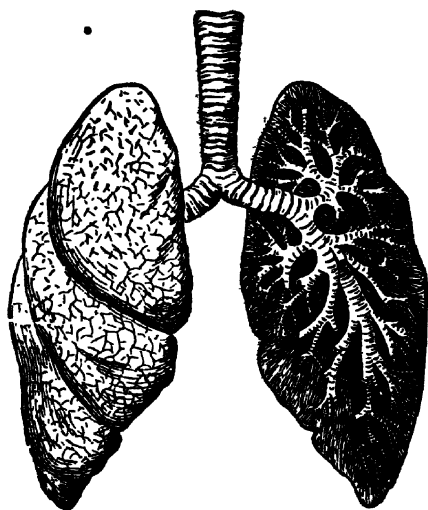


FIG. 112.—Wind-pipe and lungs. The lung on the right has been cut to show the branching tubes.

What is true of human beings is equally true of every animal which possesses lungs. They all breathe in air containing oxygen, and breathe out air containing carbon dioxide.

These facts explain why meals are required. After the muscles or the brain have been used for a certain time, a person begins to feel *hungry*. That means something is required to rebuild the used-up muscle and brain.

Food is taken, and this, after undergoing wonderful changes, supplies the blood with material which it carries off to the muscles and brain to repair them. But besides repairing waste, the food is also the source of the body's warmth. When chemical action takes place heat is developed, and it is the heat formed in this way which warms the body.

**Amount of air inhaled.**—An idea of the large amount of air essential to life may be obtained from the fact that the weight of air inhaled in the course of a day is seven or eight times greater than that of the food eaten. "The average volume of air breathed in at each breath is about 30 cubic

inches, and the volume of air which may be easily breathed in by an effort, and by expanding the chest, is about 130 cubic inches, or about four times as much. After a very full inspiration, about 230 cubic inches can be expired by a man of average height and in good health. The total capacity of the lungs, however, is much more than this—about 330 cubic inches. Thus in ordinary quiet breathing we only fill about one-tenth of the available air-space of the lungs. After expiration, a quantity of air is left in the lungs. This residual air amounts to about 100 cubic inches. An adult at rest breathes about 686,000 cubic inches of air in the course of twenty-four hours; a labourer at full work about 1,586,000 cubic inches—more than double. The amount of air passing into the lungs in a day has been estimated at 400 cubic feet in a state of rest, 600 in exercise, 1,000 in severe exertion." (Hon. Rollo Russell.)

**Purifying action of plants.**—That there is always a certain amount of carbon dioxide in the air can be proved by exposing fresh lime-water in a shallow vessel. Very soon the lime-water becomes covered with a thin white layer of chalk, which is formed by the combination of the carbon dioxide in the air with the lime in the lime-water. One reason why there is never very much carbon dioxide in the air out of doors is because there are agencies continuously at work getting rid of this gas. The chief purifying agent is the green parts of plants which occur everywhere.

When fresh watercress is put into a bottle completely full of water containing carbon dioxide in solution, and the bottle is inverted in a basin of water without allowing air to get into the bottle, it is found that, when the bottle and its contents are exposed to bright sunlight, bubbles of gas collect at the top of the bottle. These bubbles, when tested, are found to be pure oxygen. If, however, the bottle with the cress in it is kept in the dark, no bubbles of oxygen collect. Or, if a bottle of water in which carbon dioxide is dissolved be put in the sun, without any watercress, no oxygen collects in the top of the bottle.

In other words, two things are necessary for the formation of the bubbles of oxygen collected from the green plants as described. They are (1) the green vegetation, (2) the sunlight. The same conditions have been found always to hold true, thus proving that *green plants in the presence of bright sunlight have*

*the power of turning oxygen out of carbon dioxide.* They keep the carbon for themselves, and it helps them to grow. If the experiment is carefully performed, a small bunch of watercress can be proved to have increased in mass after being exposed to bright sunlight for some time, as described on p. 162.

**The balance of nature.**—The wonderful relation between animals and plants in regard to their action upon the air, is a beautiful example of the way in which nature provides for one class of its creatures even out of the refuse of another. Animals by their breathing are continually using up the oxygen and returning to the atmosphere the compound carbon dioxide—a gas which is fatal to animal life. On the other hand, plants take in the carbon dioxide, and, in the presence of sunlight, the green material in their leaves can split up the gas into the two simpler substances of which it is built, namely, carbon and oxygen. The carbon the plants keep for themselves, the oxygen they do not want, and return to the atmosphere again. In this way plants help to keep the atmosphere in a suitable condition for the healthy life of animals.

Wherever green vegetation flourishes, it is actively at work purifying the air by using up carbon dioxide and giving out oxygen. You will understand from this that open spaces and parks are of great value in towns, for they prevent the accumulation of carbon dioxide in the air as well as provide places for recreation.

It must not, however, be supposed that green plants do nothing but take in carbon dioxide and give out oxygen, for they take in a little oxygen, and in darkness they give out an appreciable quantity of carbon dioxide. But even when this is taken into account we know that vegetation is of very great use in keeping down the proportion of carbon dioxide in air. Another way in which carbon dioxide is removed from the air is by solution in the waters of the ocean, lakes, and rivers. It has been estimated that, if all the carbon dioxide produced could accumulate in the air, the proportion would be doubled in eighty-six years, and a few centuries would see the disappearance of all the higher animals.

## INCREASE OF WEIGHT WHEN CARBON IS BURNT

**177. Passing oxygen over a weighed amount of carbon.**—Place a convenient amount of charcoal in a wide, hard glass tube fitted with smaller tubes passing through india-rubber stoppers. Weigh the tube with its stoppers and contents. Arrange two U-tubes as shown; the first containing a strong solution of caustic soda, and the second lumps of the size of a pea of the same compound. Weigh the U-tubes with their contents and fittings. Join the apparatus as shown in Fig. 113. Heat the charcoal in the hard glass tube to redness, and pass oxygen slowly over it from an oxygen apparatus (p. 125), or from a gas-holder, or a cylinder of compressed oxygen. After the experiment has gone on for ten minutes, disconnect from the oxygen supply, and remove the burners. Disconnect the apparatus, and again weigh the U-tubes. When the hard glass tube is cool, weigh it again also. Notice the

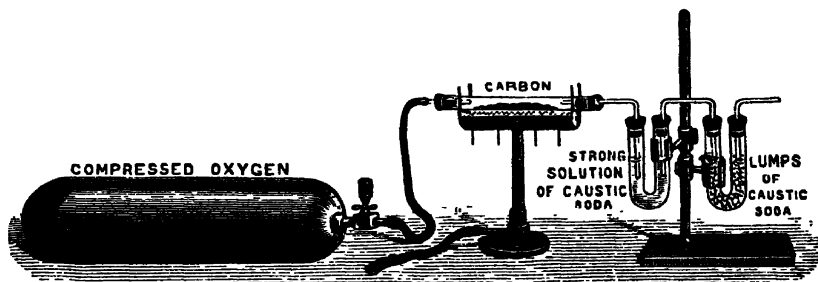


FIG. 113. — An arrangement for determining the weight of carbon dioxide produced by a known weight of carbon. The oxygen is supplied from a cylinder of compressed oxygen.

U-tubes are heavier, while the hard glass tube and its contents weigh less than before. For every gram of charcoal used up,  $3\frac{3}{4}$  grams of carbon dioxide are absorbed by the caustic soda.

**178. Increase of weight caused by burning.**—Obtain a lamp-glass like the one shown in Fig. 114. At the narrow part arrange a tray of coarse wire gauze. On the shelf so formed put pieces of caustic soda. Fit a large cork, with several holes through it, into the bottom of the lamp-glass. Through one hole in the cork insert a candle. Find the weight of the whole apparatus. Withdraw the cork, light the candle, and reinsert the cork. After the candle has burnt for a few minutes, blow out the flame, and when the apparatus is cool make another weighing. It will be found that the apparatus has increased in weight.

**When carbon burns in the air it increases in weight.**—It has now been seen sufficiently that in every case of burning the substance burnt combines with the oxygen of the air, to form a new compound. It is here necessary to study a

new change which also occurs, namely, an increase of weight which always accompanies burning. Evidence of this increase can be obtained by passing oxygen over a weighed amount of charcoal in a hard glass tube, and collecting in a suitable manner the products of the combustion which takes place. This can, in the case before us, be very conveniently done by causing the gas formed, when the charcoal burns, to pass

through a strong solution of caustic soda, a substance having the power of completely dissolving the carbon dioxide, allowing none to escape. When the experiment is carefully performed, it is always found that the increase in weight of the caustic soda is in excess of the loss in weight of the charcoal itself. Moreover, it is found uniformly that for every gram of charcoal—carbon as the chemist calls it—which is burnt,  $3\frac{3}{8}$  grams of carbon dioxide gas are dissolved by the caustic soda solution. Not only is this true of the case of burning carbon, but in every instance of burning there is a certain definite in-

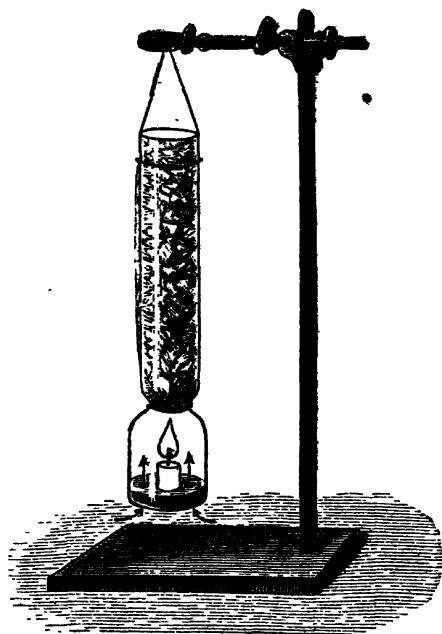


FIG 114.—Lamp-glass containing caustic soda to absorb the substances produced when a candle burns.

crease of weight, which never varies for any particular example of burning. This law of chemical combination has been known for a long time. It can be expressed in two or three ways. Either it may be said that chemical combination always takes place between definite weights of substances, or, the same chemical compound always contains the same elements combined together in the same proportions. It will now be worth while to study in detail what happens exactly when a candle burns.

**Nothing is lost when a candle burns.**—It has been seen that when iron rusts, it increases in weight. This naturally leads to the consideration of the burning of a candle. Surely, in this case, the same cannot be true. The candle gradually gets shorter and shorter, and, it will perhaps be thought, must get lighter and lighter. Of course, if a candle is weighed and then lighted, and after it has burnt for an hour is blown out and

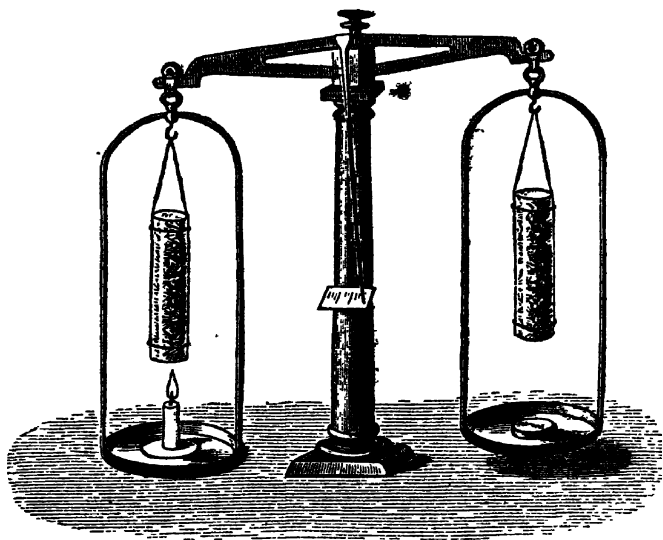


FIG. 115.—The tubes suspended from each arm of the balance contain caustic soda. When the candle in the left-hand pan burns, the substances formed are absorbed by the caustic soda above it, and their weight is shown to be greater than that of the candle used.

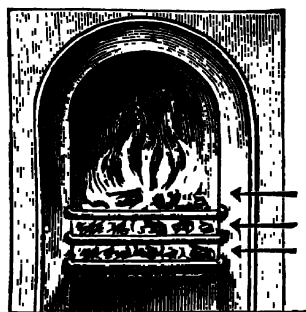
weighed again, its weight will be less than before. But this is taking no account of the compounds into which it is changed as it burns.

In the case of the rusting of iron, the rust, being a solid, stays where it is formed, and the rusty iron is easily shown to have a greater weight than the iron alone. But the candle, in taking out the active part of the air, forms two invisible gases—steam, which on a cold surface condenses into water, and the gas which turns lime-water milky,—and it is not so easy to



understand that these substances together have a *greater* weight than the part of the candle which has disappeared. Yet this is so; and the increase can be shown by simple experiments.

If an open tube, containing some lumps of caustic soda in it, is arranged above a burning candle, the smoke and the invisible gases which the candle forms in burning pass through the tube and are stopped by the caustic soda. The weight of the substances produced by the burning of a candle can thus be determined by finding the increase of weight of the caustic soda which absorbs them. If the weight of the candle at the beginning of the experiment has been found, the weight of the



Carbon monoxide uniting with oxygen to form carbon dioxide.  
Carbon dioxide and red-hot carbon form carbon monoxide.  
Carbon unites with oxygen of the air to form carbon dioxide.

FIG. 116.—Chemical changes in the burning of carbon in an ordinary coal fire.

substances absorbed by the caustic soda will be greater than the weight of the candle consumed.

**The coal fire.**—The combustion of coal is very similar to the burning of a candle. Coal is largely made up of carbon and hydrogen; its other constituents are present in smaller quantities than these two elements. When coal burns, its hydrogen combines with the oxygen of the air to form water, and its carbon with more oxygen to form carbon dioxide. The consequence is that carbon dioxide is being continually given off by fires, and under proper conditions this passes up the chimney into the outer air. In the middle of a fire, however, other changes take place. The carbon dioxide formed by the combination of the carbon of the bottom of the fire in contact with the air, in its passage through the fire comes in contact with red-

hot carbon in the centre of the fire with which it combines to form another compound of carbon with oxygen, called carbon monoxide, the properties of which are quite unlike those of carbon dioxide. Carbon monoxide is a combustible gas, and when it comes in contact with the air at the top of the fire it burns with a pale blue flame to form more carbon dioxide. Carbon monoxide gas can often be seen burning on the top of a clear fire, causing the blue flame which most persons have noticed at one time or another. Both these compounds of carbon with oxygen are injurious to life, and must on no account be breathed. It is very necessary, therefore, that there should always be a suitable means of escape for them into the open air.

Since the burning of an ordinary fire depends upon the supply of oxygen received, evidently the way to make a fire to burn is to clear the ashes from the bottom bar, and in other ways to make an easy passage for the air through the fire. It does little, if any, good to place the poker up in front of the fire, or to resort to similar superstitious devices.

On the other hand, to put out a fire we must prevent the admission of air to it. If a child has managed to get her dress on fire, it is useless to attempt to blow out the flames. Such procedure makes the fire burn more fiercely. The plan to adopt is at once to cover the fire with something—the hearthrug or some handy thick material will answer very well; this prevents oxygen getting to the fire, and it goes out very soon, because, as has been seen, oxygen is absolutely necessary in all cases of burning. This enables us to know what to do, too, when a lamp has “caught afire” owing to some defect in its



FIG. 117.—If a child's dress is afire, wrap the child immediately in a thick cloth, and the fire will be at once extinguished.

construction. Do not try to blow it out. Cover it at once with a thick duster, and cut off its supply of oxygen. If the house is on fire don't rush out leaving the doors open after you. Keep the doors and windows shut, and so try to stop the supply of air.

### SUMMARY.

**Carbon dioxide.**—When any substance containing carbon burns in the air, or in oxygen, carbon dioxide is formed. The same gas also is given off in the breath of animals. From these and other sources, large quantities of the gas are continually passing into the air.

**Breathing** gets rid of the products of a kind of burning. The oxygen taken into the lungs of an animal unites with the carbon of its body to form carbon dioxide, and with the hydrogen of its body to form water. These products are partly expelled by breathing.

**Our food** contains carbon and hydrogen; after becoming a part of the substance of the body, the carbon and hydrogen are burned, or combine with oxygen, and the products of combustion are carried away by the blood to be again removed from the body.

**Amount of air inhaled.**—In the course of a day a weight of air seven or eight times greater than that of the food eaten is inhaled by a man.

An adult at rest breathes about 686,000 cubic inches of air in twenty-four hours; a labourer at full work about 1,586,000 cubic inches.

**The green parts of plants** have power in bright sunlight to break up carbon dioxide; the carbon they keep to build up their solid parts, and the oxygen is given to the air.

**The balance of nature.**—Animals breathe *in* oxygen *from* the air and breathe *out* carbon dioxide *to* the air. Plants in bright sunshine take up carbon dioxide *from* the air, and give up oxygen *to* the air.

Carbon increases in weight when it burns in the air. One gram of carbon produces  $3\frac{3}{8}$  grams of carbon dioxide.

**A coal fire.**—At the bottom of a fire the carbon of the coal combines with the oxygen of the air to form carbon dioxide, and its hydrogen with more oxygen to form water. In passing over the red-hot coal in the middle of the fire, the carbon dioxide is converted into a compound called *carbon monoxide*, a gas which is poisonous and combustible. Carbon monoxide can often be seen burning on the top of a clear red fire with a lavender-coloured flame.

To **extinguish fires** cut off their supply of air.

### EXERCISES ON CHAPTER XIII.

1. How could you show that the same gas is breathed out by a man as is produced when a candle burns?
2. The atmosphere always contains a certain amount of carbon dioxide. How does it get there, and how would you make its presence clear to another person?

3. Describe carefully the way in which plants help to keep the atmosphere pure and sweet.
4. Breathing is connected with a particular kind of burning. What is meant by this statement?
5. Of what use are lungs to animals which possess them?
6. In what ways are plants useful to animals, and *vice versa*?
7. How would you show to a class of children that a candle does not disappear entirely when it burns, and that the substances formed as the burning continues weigh more than the part of the candle which seems to have been lost?
8. Give a short account of the changes which take place in a coal fire which is burning brightly.
9. A little girl comes running to you with her frock on fire. What should you do?
10. Why is it necessary to take food two or three times a day? When does a man want most food, and why?
11. What assistance would you give to a person whose clothes have caught fire? (1890.)

## CHAPTER XIV

### COAL GAS

**179. Distillation of coal.**—Heat some coal-dust in a hard glass tube, to which an india-rubber stopper and delivery tube are attached. After a time apply a lighted taper to the end of the delivery tube, and satisfy yourself that an inflammable gas is given off.

**180. Some properties of coal gas.**—(a) Fill a small toy balloon with coal gas, and demonstrate that the gas is lighter than air.

(b) Fill a cylinder with coal gas, as shown in Fig 121, and observe the colour and smell of the gas. As the gas can be collected over water, it is evidently insoluble in water.

**181. Products of the combustion of coal gas.**—By means of a bent glass tube and a piece of india-rubber tubing attached to the gas supply, as shown in Fig. 86, allow a small jet to burn in a cylinder. Notice the formation of drops of liquid on the sides of the cylinder. Prove that this liquid is water as in Expt. 156. Remove the gas jet. Add a little clear lime-water, and shake it up and down the cylinder. The lime-water is turned milky, showing the presence of carbon dioxide.

**Manufacture of coal gas.**—Coal gas is prepared by what is called the *destructive distillation* of coal. This process consists in heating the coal in closed retorts out of contact with the air, and collecting the products which distil over ; practically pure *coke* being left behind.

The gaseous products thus driven out of the coal pass into a receptacle where coal tar and an impure solution of ammonia gas in water, called gas liquor, are condensed. The gas bubbles through the liquid, and escapes by means of a pipe which conveys it to the *condensers*.

At the opposite end from which it entered, the gas is led from the condensers, and passes through what are known as *scrubbers*, in which the last traces of ammonia are removed. The *scrubbers* in use are of two kinds ; they are either towers filled

with coke over which water slowly trickles, or they are closed compartments provided with a revolving shaft bearing horizontal circular plates, which, being wet, expose the gas to a large surface of water. After leaving the scrubbers, the gas is conducted to a series of purifiers, where whatever carbon dioxide and sulphuretted hydrogen (a poisonous, bad-smelling compound of sulphur and hydrogen) remain are abstracted, together with some of the other sulphur compounds. The carbon dioxide should be removed, because it would lessen the illuminating power of the gas, while the sulphuretted hydrogen must be completely taken out on account

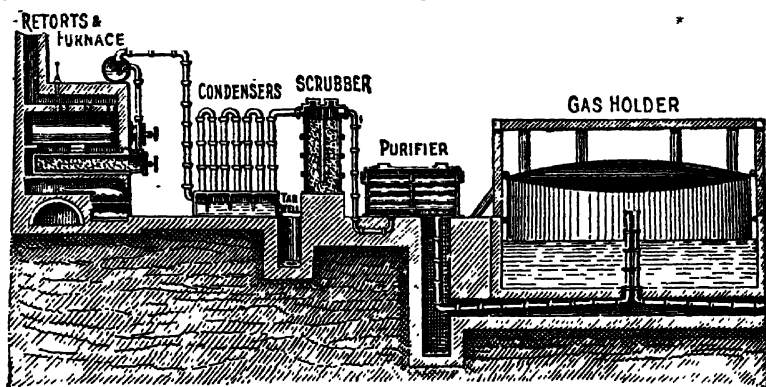


FIG. 118.—Complete arrangement for the manufacture of coal gas.

of its obnoxious smell and poisonous character. After having been measured in a suitable meter, the gas is passed into large holders called *gasometers*, from which it passes into the *mains*.

**Composition of coal gas.**—Coal gas is a mixture of several gases, which may be present in different proportions. By far the most abundant constituents are hydrogen and marsh gas (a compound of carbon and hydrogen). These, with carbon monoxide, which is present to the extent of about 6 per cent., are all combustible gases, but they none of them burn with bright flames. The three gases may be called the *combustible non-illuminating* constituents. The illuminating power of coal gas is due to the presence in it of gaseous compounds of carbon and hydrogen,

which altogether are present only to the extent of about 5 per cent.; they are therefore called the *combustible illuminating constituents*.

**Water-gas.**—The so-called “water-gas” extensively used for illuminating and heating purposes both in this country and abroad, is a mixture of hydrogen and carbon monoxide (p. 171), prepared by passing steam over highly heated hard coal. When used for lighting purposes the water-gas is first mixed with the vapours obtained by strongly heating petroleum oil.

## FLAME AND ILLUMINATION

**182. A candle flame.**—Examine the flame of a candle; notice that it is divisible into three well-marked zones. Outside is a dark yellow layer; in the middle a highly luminous layer can be made out; the interior of the flame consists of a bluish part which gives very little light.

**183. The nature of the three zones.**—(a) Take a short piece of glass tubing, about the thickness of a pencil, and open at both ends, arrange it so that one end is in the inner blue zone, while the tube itself slants away from the flame (Fig. 119). Apply a lighted match to the end of the tube away from the candle; observe that a gas which catches fire is issuing from the tube. In other words, the inner zone of a candle flame is made up of unburnt combustible gases.



FIG. 119.—Proof that the central part of a candle flame consists of unburnt combustible gases.

(b) Lower a piece of cold, thick, white note-paper on to the flame; after it has been there a second or two, raise it and examine the ring of soot which has been deposited. Round the ring of lamp-black is a second ring, where the paper is slightly charred. Repeat the experiment, bringing the piece of paper into the flame from the side, about half-way up the inner bluish zone; the results thus obtained are perhaps more satisfactory. These

results are explained by the presence of incandescent particles of carbon in the highly luminous zone. The ring of charring is caused by the outside layer, where the combustion of the vapours from the candle is complete.

**184. A gas flame.**—Examine a gas flame. Three zones similar to those in the candle flame can be made out. Using a piece of glass

tubing as before, lead some of the unburnt gases in the centre of the flame outside, and ignite them at the end of the tube. Prove the presence of particles of carbon in the middle zone by lowering a cold white plate upon the flame. Examine the deposit of soot.

**185. Result of lowering the temperature of flames.**—(a) Repeat the experiments on p. 65. Make a short coil of stout copper wire,  $\frac{1}{4}$ -inch internal diameter. Pass it over the wick of a lighted candle without touching the wick. The candle is extinguished. This is, as has been learnt before, due to the cooling effect of the wire, which conducts away the heat.

(b) Turn on, but do not light, a gas-jet. Hold over it a piece of fine wire gauze and apply a light above the gauze. Notice that the flame does not strike through. Vary the experiment by lowering a piece of cold wire gauze upon an ordinary Bunsen flame.

**186. The Bunsen burner.**—(a) Examine a Bunsen burner, or, as it is as often called, a laboratory burner. Compare its parts with Fig. 120. The metal tube, T, screws out, and reveals a small jet, J, through which the coal gas is supplied. One or two holes, A, can be closed by rotating the movable cylinder which fits round a slightly smaller fixed tube. Screw in the tube, T, open the holes, A, and connect the burner with the gas supply by means of a suitable length of india-rubber tubing. Turn on and light the gas. An almost colourless flame is obtained. You will now have no difficulty in understanding that it is a mixture of coal-gas and air which is burning. Close the holes, A, the flame immediately becomes luminous.

(b) Again obtain a non-luminous flame; lower a cold white plate into it, and observe there is no deposit of carbon. Close the holes, A, and lower the cold plate into the luminous flame you have now got; there is an abundant supply of soot. The combustion in the non-luminous flame is complete, while that in the other case is not.

(c) Shake some dust into the non-luminous flame of a Bunsen flame, and observe that it at once becomes much brighter. The dust can be suitably obtained by shaking the duster used to clean the black-board above the flame. There is a connection between the incandescent particles in a flame and its brightness.

**187. Products of combustion of a candle and coal gas.**—(a) Repeat

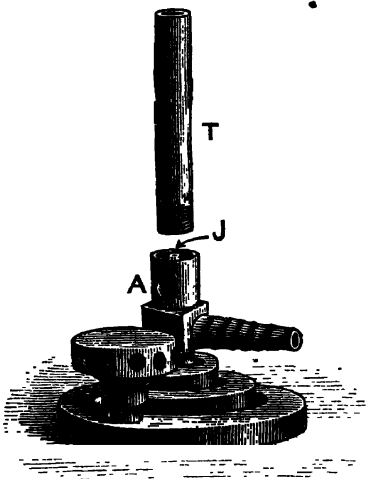


FIG. 120.—A Bunsen burner. When the flame is required to be spread out, the rose on the peg at the base of the burner is placed upon the top of the tube T.



Expts. 127 and 131, and remind yourself that in both cases of burning, water and carbon dioxide are formed.

**188. Gas explosions.**—(a) By means of the arrangement shown in Fig. 121, fill an oval soda-water bottle with a mixture of equal parts of air and coal gas. This can be done by inverting a bottle half full of water over the tube connected with the gas supply shown in the illustration. When the bottle is full of gas, close its mouth with a well-fitting india-rubber stopper or cork. Lift it out of the water, and wrap it in a duster. Holding it away from your face, pull out the cork, and apply a lighted taper. You will not fail to hear an explosion.

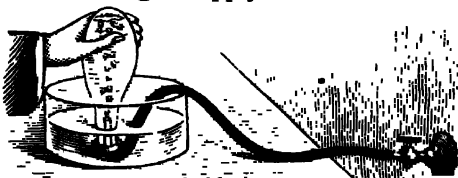


FIG. 121.—Collection of gas in a bottle.

(b) Repeat the last experiment, but this time completely fill the bottle with coal gas. On applying the burning taper in this case, the coal gas burns with a large flame, but with no explosion. In other words, while a mixture of coal gas and air is highly explosive, pure coal gas is not able to cause an explosion.

**189. The gas-stove.**—Examine a gas-stove or large ring burner like that shown in Fig. 122. The principle of its construction is the same as that of a Bunsen burner. There is always a means by which a suitable amount of air is supplied to the coal gas before it is burnt; this causes complete combustion, no deposit of soot, and a higher temperature than in the case of a luminous flame.

**190. The rate at which a gas-stove burns gas.**—It is important to be able to tell at what rate a gas-stove is consuming gas. To learn how to do this, and to form an intelligent idea of the expense attending the use of these stoves, a series of observations of the gas-meter during the time a stove is in use should be made. Put a quart of water in an ordinary kettle, place it upon the gas-stove, read the gas-meter, then turn on and light the gas. Read the meter again when the water boils, and determine how much gas has been used in the process. Similarly find out how much gas is used by the stove in an hour.

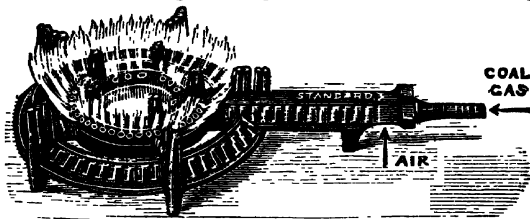


FIG. 122.—A ring burner such as occurs on the top of a gas-stove.

**The burning of a candle.**—Though a candle flame is to the majority of persons one of the most familiar of objects, it is understood

by comparatively few. Yet it is well worth intelligent study. To explain it fully, one must be acquainted with several of the most important facts of chemical and physical science. The wax

of the candle has first to be melted, and then, after the liquid wax has passed up the wick by what is called capillary attraction, it must be converted into a vapour, the constituents of which, by their combination with the oxygen of the air, burn, producing a flame. And the flame thus produced is not as simple as is first imagined. It consists of three zones each having its own particular characters. In the centre of the flame is a bluish cone of unburnt vapours, just as they were obtained by the vaporisation of the liquid wax. Outside this is the layer of the flame to which all its brightness is due ; it is impregnated by particles of carbon, which the heat of combustion has made incandescent. This can be easily proved by lowering on to the flame any cold white surface, when the particles of soot are at once deposited. The part of the flame which adjoins the air is known as the zone of complete combustion ; this layer is not highly luminous, but it has the highest temperature of any part of the flame. It is in this part of the flame, of course, that the greatest quantity of water and carbon dioxide, the products of the combustion of the candle, is formed.

It must now be explained what, from the point of view of the chemist, a flame really is. When two substances unite to form a new compound, heat is generally produced, and this heat may be sufficient to raise the temperature of the substances and the products of their combustion to such a degree that they become luminous. If, as in the case of the candle, gases are present among the products of combustion, they become luminous and pass away as flames. It may, in consequence, be stated that flames are gases produced by combustion, which have had *their temperatures raised so much that they have become self-luminous*.

If by any means the temperature is lowered, the flame is extinguished. The truth of this statement is at once demonstrated by lowering a cold coil of copper wire on to a candle without touching the wick, when the flame is put out. Or, the same thing can be demonstrated by lowering a piece of cold wire gauze on to a non-luminous flame, such as that of a laboratory burner. In this case no flame appears above the gauze, because the heat is conducted away by the wire : lowering the temperature to such a degree that flame is impossible.

**A gas flame.**—The only important difference between a gas flame and that of a candle is that the combustible body is,

in the former case, in a state of gas from the beginning, while in the latter case the solid has to be first vaporised. Just as there are three zones in a candle flame, so in the gas flame three divisions can be made out. In the centre is ordinary unconsumed coal gas, then comes the luminous part of the flame, and outside is the area of complete combustion. In both the flame of a candle and that of a gas jet the products of combustion are water and carbon dioxide.

**Forms of gas burners.**—The burners in which coal gas is used for illuminating purposes are of different shapes, so as to produce the two forms of flame in common use. These are (1) tubular flames ; (2) flat flames.

In arranging a gas burner it is necessary to make sure that the gas is being properly burnt, because it is only when this is done that the light obtained is steady, and consequently not trying to the eyesight. Attention must also be given to the

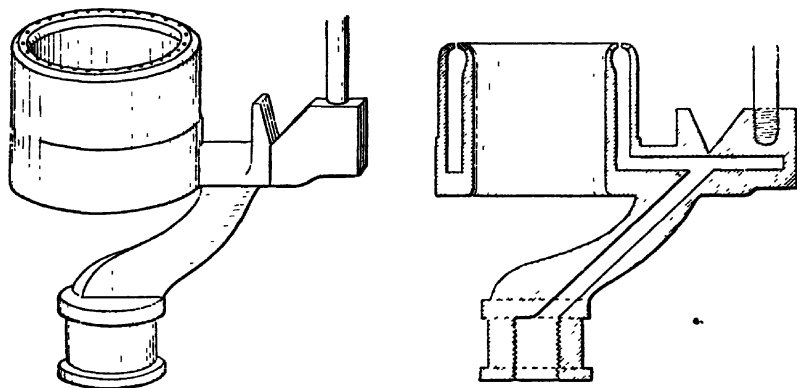


FIG. 123.—Shape and section of an Argand burner.

pressure under which the gas is supplied from the meter, and to the illuminating power of the gas. Experiments have proved that most light is obtained from a gas flame when it is on the verge of flaring.

**Argand** burners produce tubular flames. This burner consists of a tube in the form of a ring, round which a circle of small holes is bored (Fig. 123). The gas issues from these holes at a pressure slightly different from that of the atmo-

sphere. In Sugg's London Argand burner the parts where the gas is burnt are made of *steatite*, a mineral which is a bad conductor of heat.

A cylindrical chimney of glass is used with the Argand burner. This form of burner has been brought to a high state of perfection.

**Fishtail** and **Batswing** burners are designed to produce the familiar flat flames in common use. In the former, sometimes called the "union" burner, two holes pierce the surface at which the gas is burnt. These holes are inclined to one another at a right angle, the reason for this arrangement being that the collision of the two currents of gas when the tap is turned on will spread the flame out into a flat sheet.

The batswing form has a narrow slit right across the surface of the burner, instead of the two inclined holes just described. In both burners for producing flat flames the top of the burner is made of *steatite* or pottery.

**A Bunsen burner.**—The colourless, non-luminous flame of a laboratory burner is due to the more complete combustion taking place in this instance compared with what goes on in a candle flame. By a simple arrangement, which is at once understood from an examination of Fig. 120, the coal gas is, before being burnt, intimately mixed with air. The result is that the carbon of the gaseous compounds contained in the coal gas is completely converted into carbon dioxide; there is no carbon left to become incandescent, and so make the flame luminous. The combustion in both kinds of burner is of the same nature, however, water and carbon dioxide resulting in each. Since the burning in the Bunsen lamp is so much more thorough, the temperature of the flame is very much higher than that of a gas flame used for the purpose of illumination.

Gas stoves are only special forms of Bunsen burners designed to be of assistance in household economy. A perfect gas stove is one which produces the greatest heating effect with the smallest consumption of gas. Where gas stoves are used, it is of the highest importance that attention should be given to the gas meter at regular intervals, so that it may be seen at what rate the gas is being used, and so that any small leak may be detected and a waste of gas stopped.

To be effective, the supply of air to a gas stove must bear the

right proportion to the gas supplied from the main. This is the business of the maker. If this proportion is not ensured in the stove, there is a danger of the gas lighting in a wrong manner—"striking back," as it is called. When this occurs, combustion is not only incomplete, causing a deposit of soot, but an undesirable compound with an unpleasant smell, and injurious to health, is produced. Gas fires frequently "strike back" when the air-holes of the burners have become stopped with dust. When the dust is removed, the fires will burn properly.

**Products of combustion.**—It has been stated more than once that the flames of candles and of gas jets produce water and carbon dioxide. Now while the water, which at the moment of its formation is in the form of vapour, might be allowed to escape into a room with impunity, the other product of combustion, carbon dioxide, should on no account be allowed free access to living-rooms, unless suitable means for its escape are provided. In the case of gas stoves, a chimney or flue connected with the air outside the building should always be provided. This gets rid of the suffocating carbon dioxide gas ; but, unfortunately, unless intelligence is used in the way in which it is fitted, it also gets rid of a large proportion of the heat. But there is no real difficulty about the matter ; several good stoves which allow of the escape of all the products of combustion with the loss of very little heat, have been placed upon the market. A form of stove designed by Dr. F. T. Bond, which can be used for heating the air of a room without any danger arising from the carbon dioxide produced as the gas burns, is described in a later chapter.

In view of several fatal accidents which have occurred in recent years, it is worth while to point out that it is very dangerous to use "geysers" for warming water for the bath, unless the geyser is connected by a tube with the outside air, in such a way that the products of combustion can escape.

**Gas explosions.**—Pure coal gas is not explosive. In the absence of oxygen it cannot even burn. If a lighted taper is plunged into a jar of coal gas, it is found that the taper is extinguished. If it were possible, therefore, to introduce a light into a room completely full of coal gas without introducing any air whatever, the flame would simply be put out ; there would be no explosion at all.

But a mixture of air and coal gas is, like a mixture of oxygen and hydrogen, in a high degree explosive. Whenever there is any reason to fear that through an escape of gas a room has got full of a mixture of air and coal gas, it is madness to start looking for the leak with a light. Such conduct can have but one result. The proper plan to adopt is to turn the gas off at the meter, and to open windows at top and bottom, and open the doors, until the explosive mixture has had time to escape. It will then be time to think of finding the leak—a piece of detective work which is best left to a gasfitter.

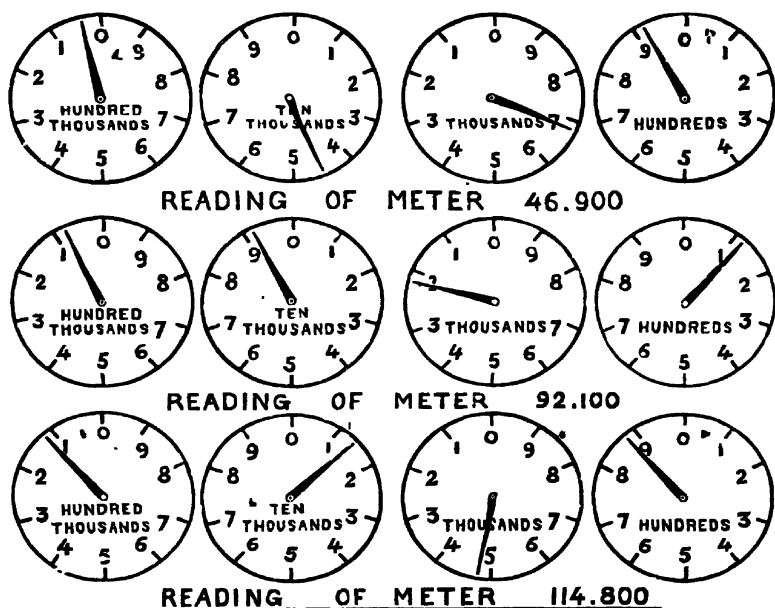


FIG. 124.—Readings of gas meter on three different dates. The dial showing cubic feet less than 100 is not shown, as it is not usually read.

**The use of the gas meter.**—To know how much gas has been used in each of the houses they supply, a gas company provides each house with a meter, and sends a man at different times to read the meter, and to record the results on a card left in the house, and in a book which he carries. The meter records how many cubic feet of gas have passed through it into the house—that is, when the meter is in proper working order.

To read a gas meter, start with the highest numbered dial, and write down one after the other the number which each hand has last passed, or which is under the hand. Add two noughts (00) to the row of figures, and this will give you the number of cubic feet indicated at the time. In Fig. 124, the appearance is shown of the dial of the meter on three different dates, at intervals of three months. In the first case the reading is 46,900 cubic feet, and three months afterwards the reading is 92,100 cubic feet, showing that, for one purpose or another, the people in the house have used 45,200 cubic feet of gas. The third time the "gas-man" read the meter, it recorded 114,800 cubic feet, indicating that during this quarter 22,700 cubic feet had been consumed.

The company charges so much for each thousand cubic feet of gas supplied, and after readings such as those described have been obtained, it is not difficult to make out the bill.

### SUMMARY.

**Coal gas** is prepared by the destructive distillation of coal. It is a mixture of gases; hydrogen and marsh gas occur most abundantly, other gases are also present, among them being some compounds of carbon and hydrogen to which the *illuminating power* of coal gas is chiefly due.

**A candle flame.**—The flame of a candle is made up of *three zones*. In the centre is a bluish zone of *unburnt vapours*; outside this is a middle zone containing *incandescent particles* of carbon; the exterior zone, the one of *complete combustion*, has the highest temperature of any part of the flame.

**A gas flame** also consists of three zones. The only difference in the two is that the candle is at first solid, while the gas has not first to be vaporised.

**Flames** are gases produced by combustion which have had their temperatures raised so much that they have become self-luminous.

**Gas burners** are of two kinds: (a) the *Argand*; (b) those producing flat flames, called the *Fishtail* and *Batwing* burners.

**A Bunsen burner.**—This burner provides a means of mixing coal gas with air before it is burnt; the result is that combustion is more complete, and a higher temperature is obtained. Gas stoves may be regarded as special kinds of Bunsen burners.

**The products of combustion** of gas stoves, gas fires, etc., must never be allowed to mix with the air of rooms. There should always be a pipe provided to lead the products of combustion to the open air.

A mixture of coal gas and air is highly explosive.

## EXERCISES ON CHAPTER XIV. •

1. Give a short description of the manufacture of coal gas.
2. What are the most important constituents of coal gas? Which of these are chiefly concerned in giving the gas its illuminating power?
3. What do you know of the structure of a candle flame? Explain what goes on in each of the zones you name.
4. Draw and describe the Bunsen burner. Explain the principle of its action.
5. How is the flame of a Bunsen burner different from that of an ordinary gas jet?
6. Describe the construction of a gas-stove. What precautions should be observed in using such stoves?
7. What is a flame? Why is a gas flame, used for lighting a room, so bright?
8. If a gas-stove is wrongly lighted, a gas with a nasty smell and very injurious to life is formed. What is the cause of this? How may it be prevented?
9. Describe some good form of gas-stove in which the products of combustion are got rid of without much loss of useful heat.
10. How are gas explosions produced? What steps should you take to prevent an explosion if you discovered there had been a large escape of gas in your sitting-room?
11. You have reason to suspect that there is a leak in one of the gas pipes in your house: state how you would go to work to find it.
12. Why is it that if a short coil of stout copper wire is passed over the wick of a lighted candle without touching the wick, the flame is put out?
13. What are the chief causes of escape of coal gas into houses? What should be done when an escape of gas is detected? (1890.)
14. In what circumstances is carbonic oxide gas (carbon monoxide) likely to be present in the air of rooms? What are its effects upon the human body? (1890.)



## CHAPTER XV

### VENTILATION AND VENTILATORS

#### PHYSICAL PRINCIPLES ON WHICH VENTILATION DEPENDS

**191. Diffusion of gases.**—*Diffusion of coal gas.*—Fill a wide-mouthed bottle with coal gas in the manner described in Expt. 188. While the mouth of the bottle is still under water, close it with a cork. Lift the bottle out of the water and place it upon a table. Take out the cork, and leave the bottle exposed to the air for some time. It will not be long before a smell of coal gas is noticed throughout the room. The coal gas has diffused throughout the air of the room.

*Diffusion of hydrogen.*—Fill a glass jar with hydrogen in the way learnt in Expt. 149. Invert the jar and hold it mouth downwards for a minute or two with its mouth uncovered. Now apply a light to the mouth of the jar. Notice there is no explosion and no flame, showing the hydrogen has all escaped from the jar. If the jar is not exposed to the air for a sufficient length of time before applying a light, there will be an explosion (p. 141), showing that the outside air has mixed with the hydrogen. *Before applying a light, therefore, take the precaution to wrap the jar in a duster.* The hydrogen escapes into the air by the process of diffusion.

*Another instance of diffusion.*—Bear in mind that ammonia gas is lighter than air, and hydrochloric acid gas is heavier than air. Take two gas cylinders of the same size, and smear their ground edges with vaseline. Into one jar pour a few drops of strong ammonia solution; into the other pour a drop or two of strong hydrochloric acid. Putting the jar containing the ammonia on the top, place the jars mouth to mouth above one another. Notice that a white cloud of the material formed by the combination of the ammonia and hydrochloric acid is produced, showing that the heavier gas has diffused upwards into the lighter gas, and the lighter gas has diffused downwards into the heavier gas.

**192. Convection currents.**—Repeat the experiments described on pp. 68-69, to explain how liquids and gases become warmed.

**193. Convection currents in a room.**—Slightly open the door of a

sitting-room, in which a fire is burning, so that a communication with the colder passage outside is made. Hold a burning candle (*a*) near the top of the door, (*b*) near the floor, and (*c*) near the handle of the door. Carefully observe what happens to the flame of the candle. It is easily seen in this way that there is a current *from the room into the passage* near the top of the door; near the floor a current *from the passage into the room* makes the flame of the candle point into the room; while in the neighbourhood of the door handle the flame is unaffected because there is no movement of the air in this place.

### The need for ventilation.

—It has been learnt that burning and similar processes are instances of the combination of some form of carbon, or some compound of carbon and hydrogen, with the oxygen of the air, resulting, when there is an abundant supply of this oxygen, in the formation of the compound, carbon dioxide, a gas in which combustion and breathing are impossible. It is true that pure fresh air always contains a certain small amount of carbon dioxide, about three parts in 10,000 of air; but this percentage has no bad effects. Indeed, air containing as much as six parts in 10,000 may be breathed with impunity. It is only when the proportion of carbon dioxide in air much exceeds this amount that the health of the person breathing it is affected. Bearing these simple facts in mind, it is not difficult to understand that it is necessary to adopt some means for keeping the air of rooms in which people have to live free from a larger amount of carbon dioxide than six parts in 10,000 of air. Ventilation, then, is the problem of managing this, without at the same time causing draughts.

**Amount of room required for healthy life.**—It has been found that an adult person breathes out three-fifths or 0·6

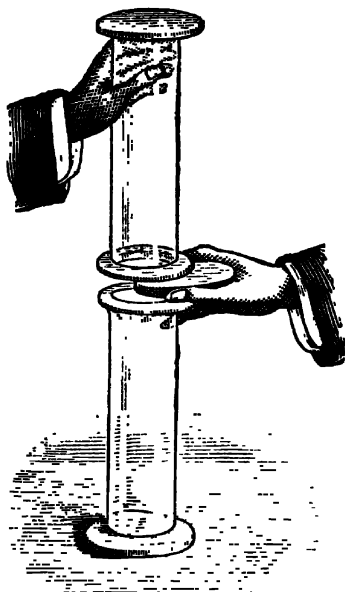


FIG. 125.—Experiment with ammonia and hydrochloric acid gas to show diffusion of gases.

of a cubic foot of carbon dioxide every hour when not in active work. Suppose a man to be in a room measuring ten feet in each of the three directions ; that is, in a room containing a thousand cubic feet of fresh air. Before this man starts breathing in the room, the air contains three-tenths, 0·3, of a cubic foot of carbon dioxide. At the end of an hour the man by his breathing has added 0·6 of a cubic foot, or altogether there is 0·9 of a cubic foot of carbon dioxide in the thousand cubic feet of air in the room, which is too great an amount to allow of healthy life.

What would be the condition of things if the room had contained two thousand cubic feet of air to begin with? When the man goes into the room the air contains twice three-tenths of a cubic foot of air ; that is,  $2 \times 0\cdot3$  cubic feet, or 0·6 cubic feet ; at the end of the hour 0·6 cubic feet of carbon dioxide is added by the man in breathing, so that altogether there is 1·2 cubic feet of carbon dioxide in the room, or 0·6 per thousand cubic feet of air—an amount which, as we have seen, is not great enough to interfere with health. It is clear, then, that to keep healthy the man must have *at least* two thousand cubic feet of air an hour, and if he is living in the room which only holds one thousand cubic feet, the air must be completely changed at the end of each half-hour. The same is true if two persons are in the room holding two thousand cubic feet of air. We can, in other words, take this amount of two thousand cubic feet of fresh air as the *smallest* amount per hour for each adult person to live healthily. It must be pointed out, however, that some authorities maintain that this allowance is too small.

**The air in a room must be changed.**—If we suppose for a moment that the air in a room can only be changed once in an hour, it is clear, in such circumstances, that, to secure two thousand cubic feet of air for every individual, each one would want a room twenty feet long, ten wide, and ten high, a size which is quite out of the question in practice. It is possible to change the air oftener than once an hour without any inconvenience in the way of draughts. In fact, in this country when cold fresh air from outside a room is allowed to enter it, it has been found by experience that the air may be changed three times in an hour and no draught ensue. If the fresh air introduced is first warmed, the change may be effected much oftener.

The following table, due to Dr. C. Childs,<sup>1</sup> gives the number of cubic feet of fresh air required per head per hour which he considers necessary for healthy life, and shows the cubic space required in cold weather with unwarmed air if changed three times in an hour, and with warmed air changed ten times an hour. Carnelley and other high authorities demand much less cubic space and air supply, practically only about one-third of the quantities given in the table.

	Cubic feet of fresh air required per head per hour.	Cubic space required in cold weather.	
		(1) With air unwarmed if changed three times in an hour.	(2) With air at 58°-60° F. changed ten times an hour.
Adult Males ... ..	3600	1200	360
Adult Females .. ...	3000	1000	300
Children ... ..	2000	666	200
Assemblages of Men, Women, and Children	3000	1000	300

The sensation of draught depends chiefly upon the speed with which the incoming air is moving, and upon its temperature. It may be said, with fair accuracy, that cold air coming into a room with a speed of less than two feet a second will not produce a draught; and, that warm air may travel into a room with a velocity of five feet a second without inconvenience.

In considering the amount of cubic space to be allowed for each person in a house, many other points must be considered. Since all fires, lamps, and candles require oxygen and produce carbon dioxide, due allowance must be made for them. It must also be taken into account how many persons are using a room at the same time before a proper conclusion as to the number of times the air in the room must be changed can be arrived at. It is also worth while to mention that children do not require so much cubic space as grown-up people.

<sup>1</sup> *Journal of the Sanitary Institute*, April, 1901, p. 26.

Mr. Francis Jones has made a detailed examination of the air of rooms lighted and heated by various means. His results show that the air of a room, however lighted and heated, is purest at the floor, less pure three feet above, and most impure at the ceiling. When a coal fire is used for heating and the electric light for lighting, the air is kept purer than by any other plan usually employed. Only by the use of this combination can carbon dioxide be kept below ten parts in 10,000 when an ordinary room has to be artificially heated and lighted. The worst air was found by Mr. Jones to be present in an open room where burners were alight on the top of a gas cooking stove, the proportion of carbon dioxide being as high as 60 parts in 10,000.

**Facts utilised in ventilation.**—Two systems of ventilation are in common use. When the problem is that of keeping the air in the rooms of an ordinary house pure and sweet, a satisfactory solution is to be found by making use of the common properties of gases. This plan is called *natural ventilation*. For the ventilation of large buildings like schools, churches, theatres, factories, and so on, *mechanical contrivances* for driving large quantities of fresh air into the space to be ventilated have to be utilised. But in this book only the former method is to be considered. Natural ventilation depends upon, first, the diffusion of gases, and second, the convection currents in gases caused by the unequal heating of different parts of a quantity of gas in communication.

Under the second heading we may include the winds, which are produced in nature in precisely the same way as the convection currents in a room in which a fire is burning.

**Diffusion of gases.**—The minute parts of which a gas is composed, "molecules" as the man of science calls them, are continually in movement. A quantity of gas, however small, will spread out until it completely fills a space, however large, in which it finds itself. Moreover, it will do this whether there is already another gas in the space or not. The result is that, whenever there are two different gases in free contact, they will soon become intimately mixed, until in every part there will be a uniform mixture of the two gases. These facts explain how it is that, if a bottle of coal gas is uncorked in a room, and allowed to remain in contact with the air of the room, the smell

of the coal gas will soon be perceived in every part of the room. In the same way the disappearance of the hydrogen from a jar of that gas held mouth downwards in a room is explained. The almost uniform composition of fresh air in different parts of the world is similarly to be accounted for.

There is a definite relation between the rate at which a gas diffuses and its density ; a light gas spreads into an adjoining space more rapidly than a heavier one, but a formal consideration of the law governing the diffusion of gases of different densities would be out of place in a little book of this character.

Owing to this power of diffusion, the air of a room rendered stuffy by the breathing of people in it, would, in the absence of all convection currents, supposing it were in free contact with the air outside, gradually diffuse into the outer purer air until all traces of impurity were lost.

Diffusion and convection must not be confused. The movements of the molecules of a fluid, which constitute diffusion, are always taking place, even when the temperature of the gas or liquid is the same throughout. The molecules move faster when the fluid is warmer, and diffusion is more rapid ; but, whatever the temperature, there is always some of this rapid movement of molecules in ordinary circumstances.

Convection currents, on the other hand, are due to the difference of temperature between two parts of a fluid in free communication. The density of the warm part becomes less as its temperature increases ; this part of the fluid consequently rises, and in due course its place is taken by the relatively heavy and cold part of the fluid.

**How convection currents are utilised.**—It has already been learnt that when air is heated it expands. The consequence is that a given weight of air when warmed fills a larger space than it did when cold ; that is to say, as its temperature is increased, its density becomes smaller and smaller. Expressed in another way, it may be said that a given volume of warm air is lighter than an equal volume of cold air. If these conclusions be applied to the air of a room in which a man is sitting at work by the side of a fire, it is not difficult to understand that, as the fire and the man warm the air, some of it near the fire rises towards the ceiling, because as it is warmed it gets lighter. A larger part of the warm air, however, passes up the chimney.

But as the air is being warmed by the man's body, the man by his breathing is rendering it impure. That is to say, warmer and impure air will be found near the ceiling. But if the air near the floor is warmed and rises, air from somewhere else must take its place, or a current of colder air is set up from the ceiling towards the floor. As will be seen later, these simple facts are attended to in all the different kinds of common ventilators in ordinary use; an exit is provided for impure air somewhere towards the upper part of a room, and an aperture for the ingress of pure air is placed near the floor.

But, as every one knows, the ventilation of an ordinary sitting-room is effected without the aid of any special apparatus. Quite enough fresh air can get into the room through the cracks which are always to be found between the window-sashes, underneath the door, and so on. Sufficient provision for the exit of impure air is provided by the chimney, especially when a fire is burning. Though these apertures ensure a satisfactory ventilation of the room, so far as keeping the air pure enough to breathe is concerned, the method is in another respect far from satisfactory, since draughts causing great inconvenience are produced. The methods of ventilation to be described immediately are all intended to avoid these draughts.

## VENTILATORS

**194. An experiment in ventilation.**—Repeat the experiment of burning a candle under a chimney glass from a lamp, and convince yourself that the introduction of the piece of card in the manner there described (p. 69) produces regular currents up one side, and down the other, of the chimney, which enable the candle to continue to burn brightly. Show the presence of the currents by the aid of the smoke from a piece of smouldering brown paper.

**195. Models of Ventilators.**—(a) Examine as many models of ventilators as are available. In each case determine the means provided for the supply of fresh air from outside, and for the escape of the impure air from the room with which the ventilator is connected.

(b) Classify the ventilators of which you have models into those used in connection with (i.) windows, (ii.) walls and chimneys, (iii.) special tubes.

**Definition of ventilation.**—The object of the different ventilators now to be described is the successful ventilation of the rooms with which they are attached. Dr. Childs, in the

paper to which reference has already been made, defines successful ventilation as that which provides a supply of air—

- (1) sufficient in quantity ;
- (2) of proper quality ;
- (3) of a suitable degree of temperature and humidity ;
- (4) properly distributed in equal proportion to each individual ;
- (5) without producing any discomfort or feeling of draught, or any gradual though imperceptible depression of health ;
- (6) with provision for complete removal of the air which has been exhaled by any individual ;
- (7) in such a way as to maintain full bodily and mental vigour ;
- (8) in such a way that the whole process is reasonably practicable and economical.

**Ventilators connected with windows.**—To avoid the draughts caused when ventilation is effected by the fresh air coming into a room by means of the cracks between window sashes, it has been found a good thing to provide a special aperture which gives the incoming air an upward tendency ; so that, before coming into contact with the persons in the room, it gets warmed to some extent. Many plans are in common use, of which the following are the most important.

**Hinckes Bird's method.**—This simple plan of giving the fresh air introduced into the room an upward tendency is easily understood by a reference to Fig. 126. The lower sash of the window is opened about three or four inches, and the opening so formed is closed with a length of board which just fills the aperture. In this way a space is left between the two sashes in the neighbourhood of the fastener, and through this the air coming in from outside is given an upward tendency.

An improved modification of this device is to make the lower sash with a double pane of glass. The principle is the same, but there is the important advantage that no light is lost.

**Louvred glass panes.**—Some authorities maintain that this expedient is too small to be really effective. The principle of its construction is the same as that of a Venetian blind. One of the panes of the upper sash, preferably one of the bottom



row, is made in strips, which can be manipulated in the same way as the bands of an ordinary Venetian blind. When open, the individual strips of glass slope upwards towards the ceiling of the room, and consequently give the air introduced from outside the necessary upward direction.

In this kind of ventilator, too, little light is intercepted (Fig. 126).

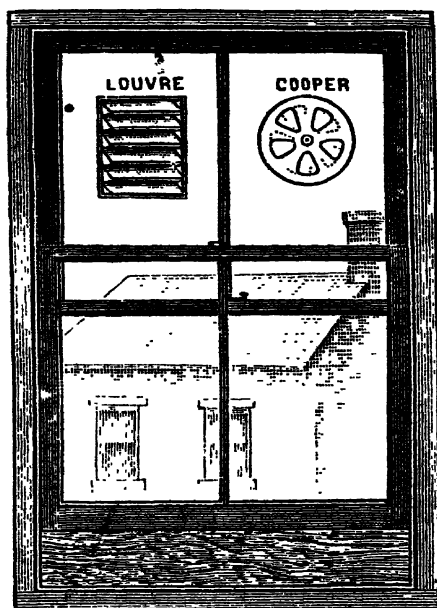


FIG. 126.—Three methods of ventilation by windows.

**Cooper's ventilator.**—As Fig. 126 makes sufficiently clear, the idea proposed by Cooper is to have a circular sheet of glass, provided with five or more holes, fixed on a window pane perforated in a similar manner, in such a way that the rotation of the circular sheet of glass brings its perforations opposite those of the window pane.

#### Other expedients.

—In schools it is a common plan to have at the top of the window a sash that swings from pivots fixed in the middle of the frame. This sash is manipulated by strings, and when open slants in

just the same way as the individual strips of a louvred glass pane.

Sometimes a number of holes are bored in the bottom horizontal bar of the upper window sash, and the air entering through these is of necessity deflected upwards. But since the holes considerably weaken the sash, the method cannot be recommended.

Double windows, which enable the air in winter to become warmed before entering the room, are really only a special case of Hinckes Bird's idea.

**Ventilators in connection with walls and chimneys.**

—The object in view when using any of the ventilators connected with walls or chimneys is the same as in the case of those fixed to windows. The chief forms are :—

**Apertures in the wall.**—A typical instance of this kind of ventilator is the well-known Sherringham valve. Into a suitably large hole in the wall fits a metal box, which is closed on the inside by a valve in the form of a heavy flap, which can be opened or shut by means of a weight attached to the end of a string passing over a pulley. When open, the flap is inclined to the face of the wall, so that the air is given the proper upward direction. On the outside the box is closed by an iron grating. The arrangement can be made out in Fig. 127.

The drawer ventilator of Stevens is very similar to a Sherringham valve. A box of the same length as the thickness of the wall where it is to be fixed, is made with a handle at one end, and minus the side

at the opposite end. It is pushed into a hole in the wall of the proper size, with the handle in the room. The consequence is that when the box is pushed from the inside of the room as far as it will go, no air can enter the room. When, however, the box is pulled out, air enters, and is deflected upwards by the end of the box carrying the handle.

**Ellison's ventilator.**—It can easily be shown experimentally that when air is forced down a cone, like a speaking trumpet, from the narrow end, it leaves the wide end of the cone with very little force. The widening and spreading of the stream of air which takes place is the cause of this. Ellison has made use of this fact in his perforated bricks (Fig. 128). These bricks are bored with holes in the shape of cones, and the bricks are built into the wall with the narrow end of the cones on the outside of the room. The air which enters the narrow end of the cone is so spread out before it enters the room that no draught at all is caused. These bricks are often placed behind the skirting boards round the room.

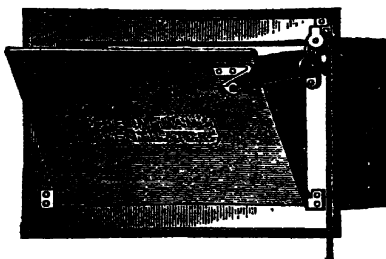


FIG. 127.—Sherringham valve for fixing in a wall to ventilate a room.

**Automatic ventilators.**—A light metal flap suitably supported so that it allows air to pass from a room, but not into it, is often fixed in the wall so that it opens into the chimney flue. The air from the room, rendered impure by the breathing of the occupants, is able to pass up the chimney, but a current from the chimney into the room at once shuts the valve so that no smoke can pass into the room. This flap is known after its inventor as Arnott's valve. Boyle's exit ventilator is of the same

kind, but, instead of a single metal flap, there are several made of mica.

**Ventilation by special tubes.**

—*Tobin's tube.*—In this arrangement the inlet for fresh air is placed near the floor, and is connected with a vertical tube not exceeding six feet in height, which is provided with a hinged cover for closing the tube when it is not in use. The plan of its construction will be understood without difficulty from Fig. 128. The advantage of Tobin's tubes is that the air, before spreading and mixing with that of the room, is warmed first of all, and then sent towards the ceiling, from which it gently descends without any draught. The tubes have the disadvantage of collecting dust, and in towns, unless special means are provided to avoid it, they introduce

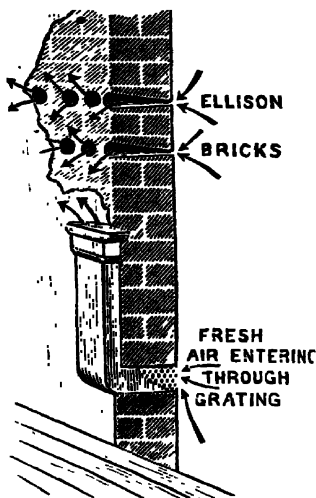


FIG. 128.—Ventilation by Ellison's bricks and Tobin's tube.

smuts and other dust into the room from the streets.

*MacKinnel's tube.*—This form of ventilator is particularly useful for one-storeyed buildings. It consists (Fig. 129) of two concentrically arranged tubes. The inner smaller tube extends from beneath the ceiling of the room, where it is provided with a circular rim, through the roof to the outside air. The outer wider tube also connects the room with the exterior air, but the opening at the lower end is on a level with the ceiling, and outside the roof it is not so high as the inner tube. The inner tube acts as an extractor of the impure warm air, its work being sometimes facilitated by lighting a gas jet in the tube. Cold,

pure air enters by the space between the tubes, and if a gas jet is burning in the inner tube this will, to some extent, warm the incoming air. The rim round the inner tube causes the current of cold air to spread out as it enters the room.

**Artificial ventilation.**—All the mechanical contrivances used in artificial ventilation have for their object one of two things. Either the im-

pure air is to be got rid of by some means of extraction, its room being taken from outside air, which enters through all available cracks and crevices ; or, the apparatus employed is designed to supply necessary quantities of fresh air from outside, and so push out the impure air, which escapes wherever it can. These

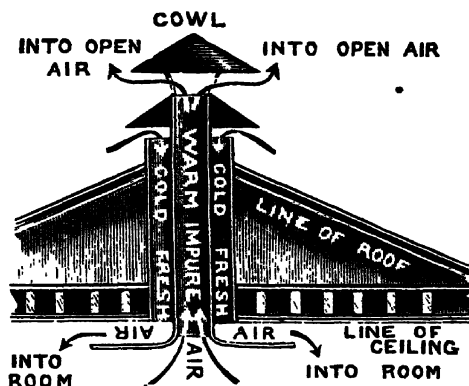


FIG. 129. — MacKinnel's tube for ventilation.

methods are only necessary in the case of large buildings, schools, theatres, and so on.

**Warming and ventilation.**—In the British Isles the temperature of the air outside a house is, in summer, as high as, or even higher than, inside. The problem of ventilation is then a simple one, for it is only necessary to throw open the windows and doors to let pure fresh air pass through the house. But during some eight months of the year the temperature in the open is much lower than that desirable for comfort in the house, and some system of warming is necessary. Most of the methods described under natural ventilation, consisting as they do of the introduction of the air from outside without any preliminary warming, are very liable to produce draughts. It becomes desirable, therefore, to consider the subject of warming buildings in connection with this question of ventilation. The simplest plan will be to describe in order the different systems of warming, and incidentally see at the same time how the problem of ventilation is modified by them.

The methods of warming in common use can be classified under three headings :—(1) Open fire-places, (2) Stoves, (3) Hot-water and steam pipes. These may with advantage be considered in order.

**Open fire-places.**—The commonest, and at the same time the pleasantest, way of warming a room is by means of the open fire-place. As ordinarily made use of, such fire-places are, however, very wasteful. It has been calculated that about seven-eighths of the heat of combustion of the fuel is wasted. In addition, the strong current produced by the fire up the chimney gives rise to cutting draughts, especially under the door. In even moderately-sized rooms, the air in the part most remote from the fire is not warmed at all, and when a good fire is burning, persons near the fire are too hot.

In studying the ways in which heat is conveyed from one place to another, it has already been learnt that such transmission can take place by conduction, by convection, and by radiation. In the case of a room with an open fire, the chief transference of heat is effected by convection and radiation. It is only in the absence of special precautions that conduction becomes an important factor ; and then it results in a considerable loss of heat by its conduction to the outer walls of the house by the iron of the grate. It has been pointed out in a previous chapter that heat is radiated through space without warming the medium through which it passes. The air of a room is not directly warmed by the heat radiated from the fire ; it is only after such heat has been absorbed by the walls, by the furniture, and by the bodies of people in the room, that these surfaces, which are thereby warmed, are able to heat the air by convection. The air near the fire and in contact with the warm fire-place is also heated by convection.

To prevent the loss of heat by conduction through the metal of the grate, a special method of construction is sometimes adopted. The plan is due to Teale. In *Teale's grate* as little metal as possible is employed. The back and sides of the fire-place are made of a non-conducting material, such as fire-brick. The brick back is inclined forwards as much as is possible, consistent with the proper escape of the smoke up the chimney, and in this way the warmth of the smoke and escaping gases is to a large extent imparted to the walls of the room. The fire-brick sides are

either parabolic in shape, or inclined towards one another. The bars in front of the fire are vertical, narrow, and close together. The ash-pit below the fire is closed by means of a metal plate called an *economiser*, which is, however, provided with an opening, which can be shut when the fuel is well alight. It will at once be evident that such a fire-place will cause the whole of the hearth and neighbourhood to get very warm, and when these grates are employed care should be taken that no combustible

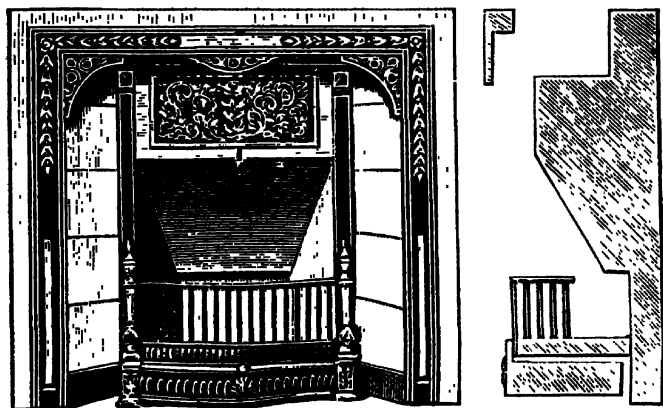


FIG. 130.—Front view and side section of a Teale's grate.

material is used in the construction of the parts of the room near the fire-place.

**Grates which warm the incoming air.**—It has been repeatedly stated that in winter, when a room requires warming, the fresh air provided in ventilating the room, if it comes directly from outside, has a temperature considerably lower than that inside the room. Unless some means are taken, therefore, to warm the air as it enters the room, draughts are almost sure to be caused. Several types of fire-grates have been devised with this object of warming the incoming air. All of them are modifications of a plan originally proposed by Cardinal Polignac, at the beginning of the eighteenth century. It will be sufficient to refer to the chief characteristics of the grate of this kind, designed by the late Sir Douglas Galton.

There is in connection with Galton's grate an air chamber,

situated at the back of the grate, and in communication on the outside of the building with the fresh air supply, and on the inside with the air of the room. The chimney flue in connection with the fire passes through this air chamber, the consequence being that the current of air from outside the room into the interior is duly warmed before its ingress. The air chamber is in communication with the room by an opening near the ceiling.

To prevent loss of heat by conduction through the metal of the fire-grate, it is lined with fire-bricks. These are arranged

in a special manner so as to effect a complete combustion of the fuel.

From what has already been learnt, it will be understood that the warm air supplied to the room by an aperture near the ceiling will as it cools get heavier and sink, and when it reaches the level of the constant current along the floor towards the fireplace,

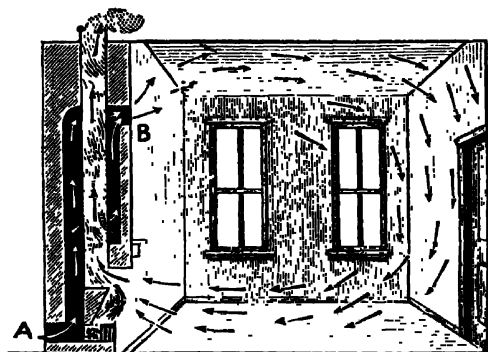


FIG. 131.—Section showing how the incoming air is warmed by the Galton grate. (Adapted from Reynolds's *Hygiene*.)

it will be absorbed by this current and be eventually carried up the chimney. During its downward passage the air is contaminated by the respiration of the inhabitants of the room. A continual circulation of air is secured by Galton's stove, a current of pure warm air inwards, and a current of impure air outwards by means of the chimney.

**Stoves.**—A stove is essentially a metal box, connected with the outer air, in which some combustible material is burnt. Its use is associated with a few advantages, but perhaps with more disadvantages. Stoves are more economical than grates, and much cleaner. They supply no radiant heat to speak of, and the consequence is that the occupants of dwellings warmed by stoves are warmed chiefly by the air, which is raised to a high temperature by convection by direct contact with the stove.

Stoves are commonly made of cast-iron, and this iron often gets red-hot. This fact leads to several dangers. Red-hot cast-

iron is very porous to the gases formed by the combustion of the fuel. Among these products in a stove in which coal or coke is burnt, are carbon dioxide and carbon monoxide, both of them fatal to life, the latter gas being absolutely poisonous. Because of this porosity of highly heated cast-iron, stoves should always be lined with fire-brick ; they may with advantage be covered on the outside with tiles.

A common complaint against stoves is that they give rise to an uncomfortable feeling of dryness in the air of the room in which they are used. Many people contend, indeed, that this dryness is not only unpleasant but harmful. Mr. Thomas Fletcher maintains that this dryness is due, in the case of gas stoves, to the escape of some of the products of combustion into the room. The feeling of dryness in the throat and nostrils is caused, Mr. Fletcher says, by the compounds of sulphur which find their way into the room with escaping products of combustion. Other authorities are of opinion that the so-called dry feeling is to be traced to the burning of tiny particles of organic matter present in the air, by their contact with the highly heated metal of the stove.

A simple method of mitigating, to some extent at least, this feeling of dryness is to have a bowl of water exposed to the air of the room. The continual evaporation which will go on will supply any deficiency of water vapour in the air which may result from the higher temperature to which it is heated ; and it will also serve, to a large extent, the purpose of dissolving any gaseous compounds of sulphur which may be present in the air of the room.

Whatever the material which is burnt in a room, be it coal, coke, or gas, there should always be a flue in connection with the open air to conduct away the products of combustion.

Stoves easily lend themselves to the purpose of at the same time warming and ventilating a room. The plan adopted in the case of stoves in which coal or coke is burnt can be made out from an examination of Fig. 132, which shows the " Sandringham " warm air stove, in which smokeless coal may be burnt. Fresh air is provided by means of a tube under the floor in connection with an opening at the bottom of the stove. This opening is in direct communication with a chamber of air forming the outer part of the stove, in contact, by means of a grating,



with the air of the room. The products of combustion from the stove, as well as the impure air of the room, pass out by way of the flue. The fresh air from outside is in this stove effectually warmed before it enters the room.

A gas stove which brings about the same result in a very effective manner is that designed by Dr. Bond, and called the "Euthermic ventilating gas stove." The fresh air is provided through the tube, E (Fig. 133), into the air chamber, DBC, in

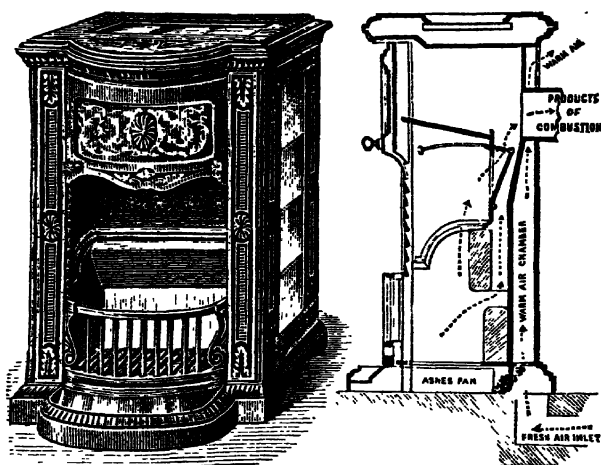


FIG. 132.—"Sandringham" warm air stove.

communication with the air of the room through the grating, F. The products of combustion from the burners, shown at K, pass through the space between the inner air chamber and the outer walls of the stove, and escape by way of the flue, H.

**Heating by hot-water and steam.**—The methods of heating by hot-water and steam are more useful for large institutions than for dwelling-houses. The commonest way of utilizing hot water is that known as the *low pressure system*. The principle underlying this method is the same as that exemplified in the experiment on p. 69, where convection currents are shown in the apparatus made of glass tubes in connection with a reservoir of water.

The low pressure system, reduced to its simplest terms, is

shown in Fig. 134. The boiler, B, the pipes, *ab*, *ec*, and the coils, C, C, are nearly full of water. This water, before the fire in the stove below B is lighted, is cold. When the layers of water near the bottom of the boiler are heated, they get lighter and rise up the tube, *ab*, to the top of the building, where the water gives out its heat to the upper room. The place of the rising water is taken by cold water from the other pipe, *ecd*, which, in its turn, gets warmed and rises. In the meanwhile, the warm water in the upper rooms has given out its heat, and becomes cold again, sinking to the level of the boiler to be again warmed. This circulation continues so long as there is a fire in the furnace, and in consequence a supply of heat can be kept up as long as is desired. The water in the pipes of a low pressure system never gets hotter than a few degrees below the boiling point of water, a temperature which does not interfere with the healthy life of persons living in the building.

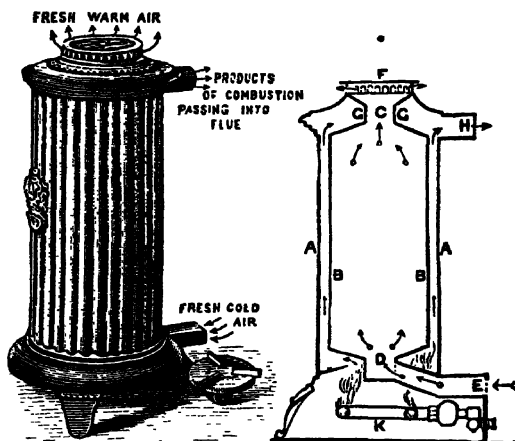


FIG. 133.—Bond's Euthermic ventilating gas stove.

This plan for warming rooms of a large size is on the whole a very satisfactory one. It is under good control, and the coils of pipes from which the heat is given out can be arranged anywhere that is thought desirable.

Another hot-water system is in use under the name of the *high pressure system*. Its use cannot be recommended for ordinary purposes, since it requires constant skilled attention. In the high pressure system much smaller pipes than those in the low pressure plan are used. Instead of a boiler of the ordinary form, a coil of iron tubing is used, and this is placed in the fire of the furnace.

The small amount of water contained by the pipes is soon heated to as high a temperature as is ever attained. There is

perhaps more risk of an explosion in this system than in the low pressure, but in the case of such accident there is no great danger, as when a pipe breaks there is merely an escape of steam and water and no pieces of metal fly about.

Steam is generally only employed in special circumstances.

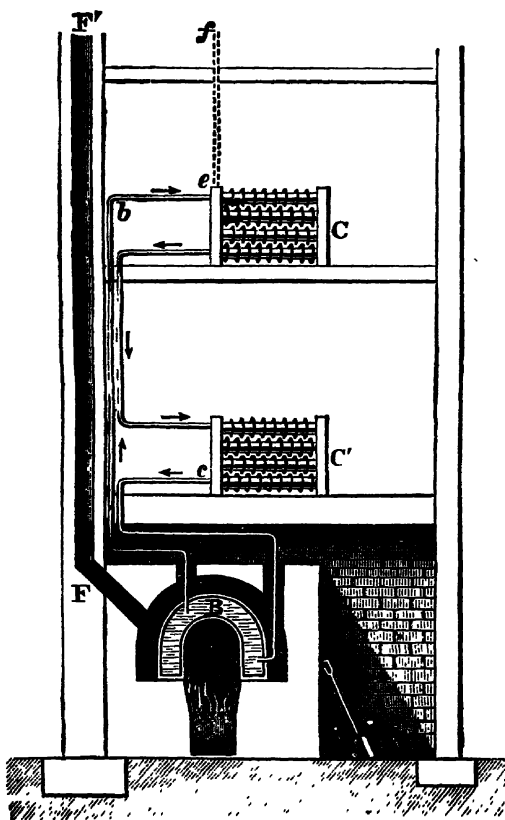


FIG. 134.—Arrangement of pipes for heating a building with hot water.

If steam is already being employed for the purpose of driving machinery, it is convenient to use it also for warming purposes. Consequently, in places like laundries it is a convenient source of heat. In recent years it has become customary to warm railway carriages by steam.

## SUMMARY.

**The need for ventilation.**—*Breathing* and *burning* continually increase the amount of carbon dioxide in the air. For healthy life, air containing more than six parts of carbon dioxide in ten thousand should not be breathed. Ventilation is the problem of keeping the percentage of carbon dioxide below this amount without causing draughts.

**Space necessary for health.**—To keep healthy, a man should have at least two thousand cubic feet of fresh air an hour.

**Changing the air of a room.**—The sensation of draught depends chiefly upon the *speed* with which fresh air enters a room. If the speed of incoming air is *less than three feet a second*, no draught is caused. If air is *warm*, its speed may be *five feet a second*.

**Two systems of ventilation.**—*Natural* ventilation makes use of the common properties of gases, such as *convection currents*, and *diffusion*. For large buildings *mechanical contrivances* are employed.

**Diffusion** of gases is due to the incessant movements of the molecules of a gas. A light gas diffuses more rapidly than a heavy one.

**Convection currents in gases** have been studied sufficiently in Chapter VI.

**Ventilators** may be divided into those in connection with (a) windows, (b) walls and chimneys, (c) special tubes.

**Window ventilators.**—Hinckes Bird's method, louvred glass panes, Cooper's ventilator, and others may be mentioned under this class.

**Wall and chimney ventilators.**—The Sherringham valve, Stevens's drawer ventilator, Ellison's perforated bricks, Arnott's and Boyle's valves, may be classed in this way.

**Special tubes.**—Tobin's and MacKinnel's are typical instances.

**Methods of warming.**—These may be classified under three headings:—(1) Open fire-places, (2) Stoves, (3) Hot-water and steam pipes.

**Open fire-places.**—These are wasteful and produce draughts. Teale's grate diminishes loss of heat by conduction and ensures the slow combustion of the fuel. Galton's grate is specially designed to warm the incoming air; it combines ventilation and warming.

**Stoves.**—These are more economical than grates, and much cleaner. Since red-hot cast-iron is porous to the products of combustion, stoves should be lined with fire-brick, and be covered outside with tiles. There should always be a flue in connection with the open air, to carry away the products of combustion. Bond's stove serves both to warm and ventilate a room.

**Hot-water and steam pipes.**—The *low pressure system* of hot-water pipes is the commonest. It is under good control, and is on the whole very satisfactory. The *high pressure system* requires constant attention; with it much smaller pipes are used, and the small amount of water they hold is soon heated to a high temperature.

If *steam* is already in use for driving machinery, it can also be used conveniently for warming purposes.

## EXERCISES ON CHAPTER XV.

1. What do you understand by the "diffusion of gases"? Describe an experiment you would perform to make your meaning clear.

2. Name the different ways in which heat may pass from one place to another. Which of these is made use of when a room is warmed by an open fire?

3. Why is it necessary to ventilate rooms in which people live? Explain when a room is properly ventilated.

4. How often should the air of a room be changed? Make it quite clear in your answer that you understand what circumstances must be taken into account.

5. How are convection currents made use of in ventilating a room? Explain when a draught is likely to be caused by utilising them.

6. It is very common to feel a draught when sitting between an open fire-place and the door, especially a draught to one's legs. Why is this, and how may it be prevented?

7. Draw and describe two kinds of ventilators used in connection with windows. Point out the advantages and disadvantages of the forms you select for description.

8. Draw a Sherringham valve and one of Ellison's perforated bricks. What is the object of each device?

9. Describe a simple form of automatic ventilator. With the aid of a sketch explain its action.

10. Enumerate the advantages of a Tobin's tube as a ventilator for a school class-room. Has it any disadvantages?

11. Sketch MacKinnel's tube ventilator, and explain the way in which it serves to keep a room ventilated.

12. What is meant by a slow combustion grate? Describe any form with which you are acquainted.

13. What are the advantages of cooking by gas? What conditions should a gas-cooking oven fulfil? (1888.)

14. What are the forces which produce natural ventilation? How may the wind be practically utilised? (1888.)

15. Describe in detail a grate or stove provided with an arrangement for the introduction of warm, fresh air. (1889.)

16. Explain the principle of construction of an ordinary fire-grate, and state its advantages and disadvantages. (1892.)

17. Describe and compare the action of ordinary fire-places and hot-water pipes in warming and ventilating rooms. (1894.)

18. What are the best means of ventilating a room without causing draughts? If possible, illustrate your answer by a sketch. (1894.)

19. What is carbonic acid? What are its sources? What part does it play as a sign of good or bad ventilation? (1896.)

20. What are the objections to stoves as a means of heating dwelling-rooms? Describe any two forms with which you may be familiar, and point out their respective advantages and disadvantages. (1896.)

## CHAPTER XVI

### THE CHIEF CONSTITUENTS OF FLOUR

#### GLUTEN

**196. The appearance of good flour.**—Examine a sample of ordinary flour, such as is used for bread-making. Its colour is white, with just a suggestion of cream colour. It has a faint and pleasant characteristic odour, and it is sweet and somewhat “nutty” to the taste. Rub a little between the finger and thumb, and notice the “feel.” Press some in the hand. The flour sticks together slightly, but, unless it is too damp, it falls apart again when the pressure is removed.

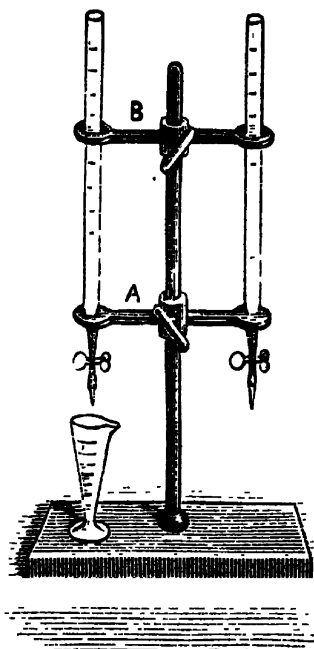


FIG. 135.—Burettes supported in a burette stand.

**197. Dough.**—Weigh out 10 grams of the flour into an evaporating basin, and add water from a burette (Fig. 135) in small quantities at a time, stirring with a strong glass rod until a uniform dough is obtained, which can be kneaded between the fingers without either feeling sticky or leaving particles of flour behind. This point can be determined very exactly; and it will be found that a certain weight of flour from the same sample always requires the same amount of water to make a dough of this consistency. The greater the quantity of water required, the “stronger” is the flour. If different qualities of flour can be obtained, find how much water is required in each case to form a dough

with 10 grams; and notice that the samples which require most water are the strongest when the dough is pulled gently by the fingers.

**198. The separation of dough into its constituents.**—Tie up the dough in a piece of muslin or an old handkerchief, and knead it gently between

the fingers in a basin of water. The water at once becomes milky, owing to the starch which comes from the dough. Continue the kneading and change the water from time to time until no more milkiness is produced. Now open the cloth and examine the gluten which is left.

**199. The properties of wet gluten.**—Freshly prepared gluten of good quality is greyish-yellow in colour. From the method of its preparation it is evident that it is insoluble in water. When pulled it is found to be very tenacious and elastic,—for it springs back when released, like thin india-rubber. It is practically free from taste and smell. So long

as it is wet it does not stick to the fingers, but if allowed to dry slightly it becomes sticky.

**200. The proportion of dry gluten contained in flour.**—

Weigh a clean and dry flat-bottomed dish and spread out the gluten as evenly as possible in it. Put the dish into the air-oven (Fig. 136), and heat it at a temperature of just over  $100^{\circ}\text{C}$ . for two or three hours. Then cool the dish and weigh again. The difference between this weight and that of the empty dish is the weight of the dry gluten obtained from the 10 grams of flour taken (Expt. 197). Multiply this by 10 to obtain the percentage of dry gluten in the sample of flour.

Good wheaten flour contains from 10 to 13 per cent. of dry gluten.

**201. The properties of dry gluten.**—The gluten has changed in colour during drying, and has become very hard and brittle, like glue, and difficult to remove from the dish. Break a piece off, and soak it in water. It will be found impossible to restore the gluten to its former plastic condition.

**202. The effects of heat upon gluten.**—Heat the dish containing the gluten by means of the Bunsen flame. The gluten blackens (showing the presence of *carbon*), shrivels up, melts, and gives off smoke with an odour somewhat like that of burning hair or feathers. This peculiar smell indicates the presence of *nitrogen*. At last the gluten bursts into flame, and if the heating is continued, burns away.

**203. Gluten readily decomposes.**—Prepare another piece of gluten, and, without drying, put it aside for a few days in a dish. It soon putrefies and acquires a most offensive smell.

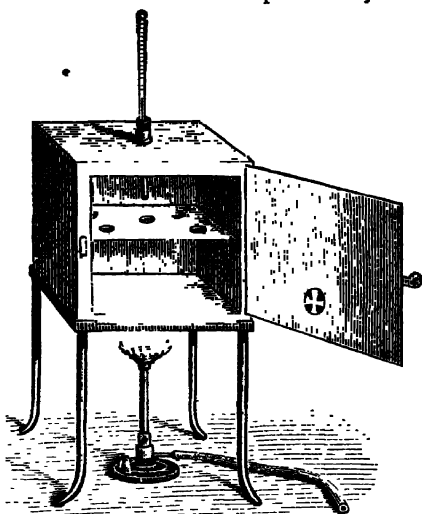


FIG. 136.—Air oven heated by rose-burner.

**Wheat flour.**—Wheat flour is a white or creamy powder,

with a faint pleasant odour. It has a "nutty" and slightly sweet taste. When good flour is taken up in the hand and squeezed, it shows a slight tendency to stick together, but it readily falls apart again when the pressure is removed.

**Dough.**—When flour is mixed with water, it forms a plastic mass, known as *dough*. Some flours require more water than others to form a dough of a given consistency. These are known as "strong" flours. The dough obtained from them is very tenacious.

**The chief constituents of flour.**—When dough is tied up in a piece of muslin and kneaded in water, it separates into two constituents with very different properties. One of these, *starch*, comes through the meshes of the cloth and turns the water milky. If allowed to settle, it forms a white deposit, which may be collected and dried. When the starch has been completely separated, the muslin is found to contain a greyish indiarubbery mass, called *gluten*.

**Gluten.**—Fine wheaten flour contains from 10 to 13 per cent. of gluten. This, when freshly separated from dough, is a tenacious, elastic substance, somewhat of the consistency of thin india-rubber.

When gluten is dried at the temperature of boiling water, it becomes very hard and brittle, like glue, and cannot be again transformed into the plastic condition. When it is strongly heated, it blackens, shrivels up, and smokes, giving off a strong odour resembling that of burning feathers or hair, due to certain compounds of nitrogen. On continued ignition it catches fire, and at last burns completely away.

When gluten in the plastic or indiarubbery state is exposed to the air, it soon begins to decompose or putrefy, and gives off gases which smell most offensively.

Gluten has the following percentage composition, that is, in 100 parts of gluten the following elements are present in the proportions shown :

Carbon . . . . .	52.6
Hydrogen, . . . . .	7.0
Nitrogen . . . . .	16.0
Oxygen, with traces of sulphur . . . . .	24.4
	<hr/> 100.0



**Proteids.**—The composition of gluten is almost exactly the same as that of *albumin* or white of egg, in spite of the great differences in their appearance. In fact, both gluten and albumin are included by chemists in the group of bodies called *proteids*. Later on we shall have to consider several other proteids. The great importance of these compounds lies in the fact that animals cannot live unless a certain amount of proteid matter is included in their food; for this is the only form in which the absolutely necessary element *nitrogen* can be made use of by the animal body.

**The value of gluten as a food.**—Gluten is not only a proteid, but one which is easily digested. On this account it has a very high nutritive value.

Gluten itself, mixed with only a very small quantity of starch, is sometimes baked into loaves and biscuits for the use of patients suffering from diabetes. The starch—which is considered injurious in cases of diabetes—is removed before baking by kneading the dough under a stream of water. Bread made in this way is known as *gluten bread*.

## STARCH

**204. Appearance of starch.**—Examine some lumps of ordinary laundry starch. The masses have a general resemblance to irregular prisms, which are often twisted. Carefully distinguish these from crystals.

Starch is non-crystalline. The colour is a dead white. The pieces of starch are very brittle and may easily be reduced to a fine powder.

Crush a lump in the mortar, and rub a little of the powder between the finger and thumb.

The starch has a somewhat “harsh” feel. It is quite devoid of taste and smell.

**205. Make a solution of iodine.**—Add a small crystal of potassium iodide to about a cupful of water, and then add iodine in *small* particles at a time, stirring continually, until the solution is the colour of sherry. Put the solution into a bottle and label it.

Iodine dissolves much more freely in water containing potassium iodide than in water alone.

**206. The action of iodine upon starch.**—Pour a drop of the iodine solution on a lump of starch. The starch at once becomes dark blue in colour. *The production of this colour on adding iodine solution is the best test for starch.*

**207. Starch is present in potatoes.**—Cut a slice from a potato, and pour a drop of iodine solution on it. In a few moments the surface of the potato is seen to be covered with tiny blue dots.

**208. Starch is present in grains of wheat and maize.**—(a) Soak grains of wheat and maize in water over-night, and then cut them through with a sharp knife, and pour iodine solution upon the cut surfaces. The blue colour shows the presence of starch.

(b) Try the same experiment with rice, sago, and tapioca. Prove the presence of starch in flour.

**209. Starch is formed in green leaves.**—Take a few leaves from a plant (a fuchsia answers very well) which has been exposed to sunlight for some hours; and put them into boiling water for a few minutes. Then place them in methylated spirit until they are bleached. Now wash them by placing in water; pour the water off, and cover them with iodine solution. The leaves turn blue, showing the presence of starch.

**210. The microscopic appearance of starch grains.**—(a) Put a small drop of water upon a clean microscope-slide. Scrape very gently the

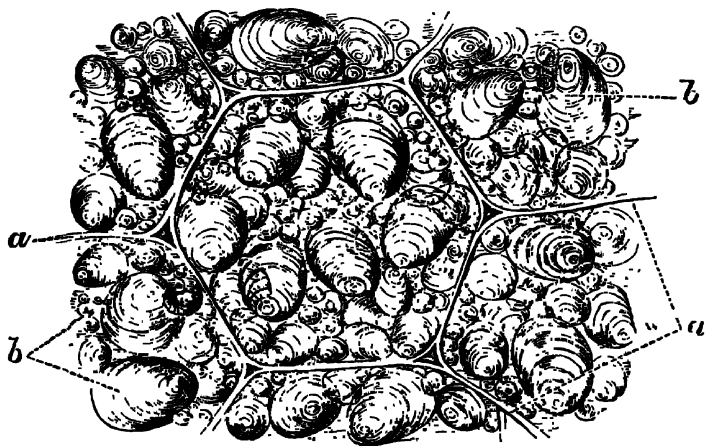


FIG. 137.—A thin slice of potato, highly magnified. *a*, cell-walls, composed of cellulose; *b*, starch grains.

cut surface of a potato with the point of a knife, and then dip the knife-point into the drop of water. This becomes milky. Now gently lower a clean "cover-glass" upon the drop, being careful not to enclose any air-bubbles. Put the slide upon the stage of the microscope and examine with the low power (the "1-inch" objective). Notice the large number of glistening particles (Fig. 137). These are the starch grains of the potato. They vary much in size and shape, the smallest being almost spherical, while the largest are roughly ovoid. By careful focussing, the large grains are seen to be made up of a number of concentric layers surrounding a dark spot, which is generally near the smaller end of the grain.

(b) Place a drop of iodine solution at one side of the cover-glass, taking

care that none gets on its upper surface, and put a small piece of blotting-paper at the opposite edge of the cover-glass. This draws the iodine solution under the cover-glass. Notice that the grains become blue as the iodine solution reaches them. Potato starch is good for practice because the grains are so large.

(c) For the examination of most other starches a high power (the " $\frac{1}{4}$ -inch" objective) is necessary. Add a particle of "corn-flour" (usually maize flour) to a small drop of water so as to turn it just milky. Examine this with the high power in the same way. In this case the grains are of a quite different shape.

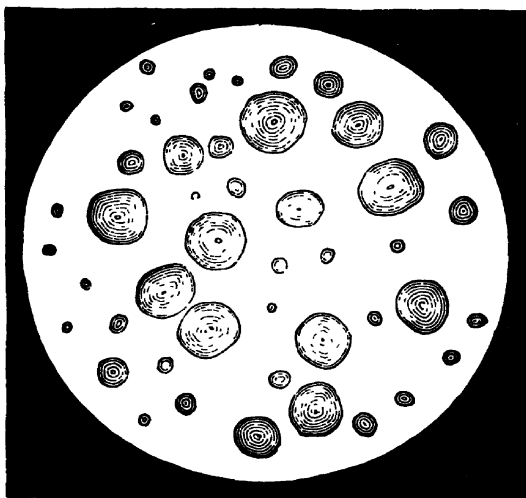


FIG 138.—Grains of wheat starch, highly magnified.

(d) Similarly, examine wheat flour (Fig. 138) and rice-starch (Fig. 139) and notice the differences in the size and shape of the starch grains.

**211. Starch is insoluble in cold water.**—Mix with cold water as much starch as will go on a shilling; filter carefully so that the liquid which passes through is *quite clear*. Add iodine solution to this. The liquid does not become blue showing that no starch has been dissolved.

**212. Starch is soluble in hot water.**—(a) Take the same quantity of starch as in the last experiment and make it into a thin cream with a little cold water in an evaporating-dish. Now add about a cupful of boiling water and stir. The starch swells up and forms an almost clear solution. Take a drop of this and examine it, after covering with a cover-glass, under the microscope. No starch grains can be seen. They have been burst by the boiling water, and most of their contents have dissolved.

(b) Add a drop or two of the starch solution (often called *starch paste*)

to a cupful of water, and add iodine solution. Notice that, even with the very small quantity of starch present, a beautiful blue colour is formed, showing that the test is a very delicate one.

Heat the liquid; the blue colour is destroyed. Cool the liquid again; the colour reappears.

**213. Starch-paste is neutral.**—(a) With a glass rod place a drop of

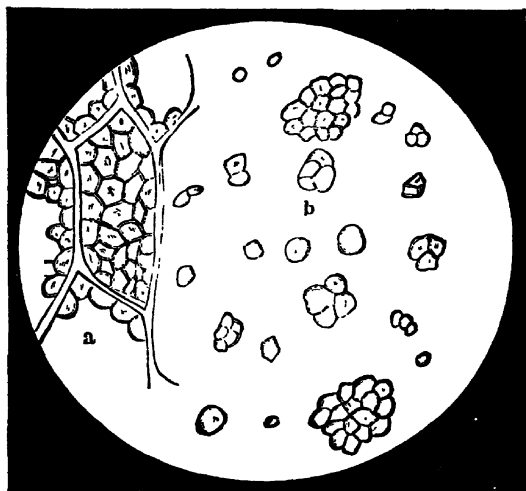


FIG. 139.—Grains of rice starch, highly magnified.

starch-paste upon (1) red litmus paper, and (2) blue litmus paper. In neither case does the litmus paper change colour. Starch paste is neutral.

(b) Similarly, place drops of vinegar and weak sodium carbonate solution upon red and blue litmus papers. Observe what occurs, and tabulate the results as follows:—

Substance.	Action upon red litmus paper.	Action upon blue litmus paper.	Acid or alkaline.
Vinegar . .	None.	Paper turned red.	Acid.
Sodium carbonate . . . .	Paper turned blue.	None.	Alkaline.
Starch paste .	None.	None.	Neutral.

**214. Starch paste will not pass through moist parchment paper.—**

(a) Take a thistle-funnel, and with a file cut through the stem about six inches below the head. Wet a piece of parchment paper (having previously held the paper up to the light to be sure there are no holes in it), and tie it tightly across the mouth of the funnel. Fill the head and about an inch of the stem of the funnel with thin starch paste. This can easily be done by means of a pipette (Fig. 64), or of a "canula" such as is used for filling fountain-pens. Now put the funnel into a beaker of water, in the manner shown in Fig. 140, and put the arrangement aside for a few hours. After that time add iodine solution to the water in the beaker. No blue colour is formed, showing that no starch has passed through the membrane.

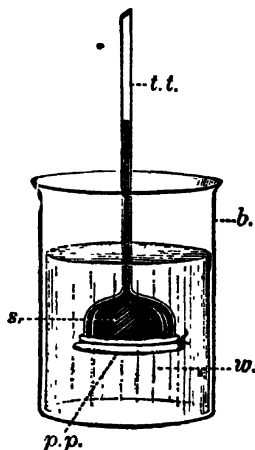


FIG. 140.—A simple dialyser. *b.*, beaker; *w.*, water; *p.p.*, parchment paper; *t.t.*, stem of thistle-funnel; *s.*, starch paste in head of thistle-funnel. The starch paste does not pass through the parchment paper, but a solution of sugar would.

(b) The funnel and the parchment paper form what is called a *dialyser*. A substance which, when dissolved in water, is unable to pass through a moist membrane such as parchment paper is called a *colloid*. Starch is one of the most perfect colloids known.

**215. Glucose will pass through moist parchment paper.**—Repeat Expt. 214, but instead of starch paste use "glucose" (a variety of sugar used for making syrups, etc.). To show that some of the glucose has passed through the dialyser, take about a teaspoonful of the water in the beaker, and put it in a test-tube with twice as much Fehling's solution.<sup>1</sup> Put the test-tube into a vessel of boiling water, and notice that in a short time the blue colour of the solution disappears, and the liquid becomes turbid with a red precipitate.<sup>2</sup> This shows the presence of glucose.

A substance which, like glucose, will pass in solution through a moist membrane, is known as a *crystalloid*.

**216. The action of sulphuric acid upon starch.**—Make about half a pint of thin starch paste, and add to it about 20 drops of sulphuric acid which has been diluted with five

times its volume of water. Put the mixture into a flask, and keep it for some time at a temperature just below the boiling point. From time to time take out a few drops with a pipette (Fig. 64), and test with iodine solution. As the heating is continued, the colour produced with iodine changes, until it is at last reddish brown. If now a few drops be allowed to fall into alcohol, flocculent matter will separate.

<sup>1</sup> Prepared by adding to a solution of copper sulphate first tartaric acid and then caustic soda until the blue mixture is clear.

<sup>2</sup> The test-tube may be readily cleaned afterwards by means of a little ammonia solution.

This substance is *dextrin*, sometimes known as British gum.

At this point raise the mixture in the flask to boiling for about fifteen minutes, and test again with iodine. Very little colouration is now produced. Cool the liquid, and place some of it in a test-tube with Fehling's solution. On heating the test-tube in boiling water, the liquid becomes red and turbid, showing the presence of *glucose*.

**217. The action of malt extract upon starch.**—To about half a pint of thin starch paste add a tablespoonful of malt extract, and set the mixture aside for a few hours. Then test a portion of it with iodine solution. No blue colour is produced, showing that the starch has been changed into something else. Test another portion with Fehling's solution. The red precipitate which is produced indicates that the starch has become changed into a sugar.

In this case the sugar is not *glucose*, but *maltose*.

**218. The action of saliva upon starch.**—Put a few drops of starch paste into a test-tube and cover it with saliva. Put in a warm place for twenty minutes. Then add water and test for starch and sugar as in Expt. 217.

**219. How to change starch into dextrin.**—Put a small piece of starch into a porcelain dish, and moisten it with very dilute nitric acid. Put the dish into the air-oven, and heat the oven with the Bunsen, using a rose-burner to spread the flame. Regulate the gas-flame so that the temperature of the oven, as indicated by the thermometer, is kept at about  $150^{\circ}$  C. for some time.

The starch changes into a glassy substance, colourless or yellowish, which is soluble in water; the solution gives neither a red precipitate with Fehling's solution, nor a blue colour with iodine solution.

This glassy substance is *dextrin* or British gum.

**220. The effects of heat upon starch.**—Heat a piece of starch in a porcelain crucible by means of the Bunsen flame (Fig. 87). It turns yellow, brown, then black (owing to the separation of *carbon*), and at last catches fire and burns. Upon prolonged heating, it burns almost completely away.

**221. Experiments with substances allied to starch.**—(a) *The effects of heat upon filter paper.*—Cotton-wool and the best filter paper consist of almost pure *cellulose*, a substance similar in percentage composition to starch. Roll up a piece of filter paper, and put it in a porcelain crucible and heat with the Bunsen. Only a very small quantity of light ash is left.

(b) *The effects of heat upon wood.*—Weigh two or three wooden matches, after cutting off the heads. Break them into small pieces, and heat them strongly in a previously weighed porcelain crucible. When nothing but a white ash is left, cool the crucible and weigh again. Calculate the percentage of ash or mineral matter.

Wood is largely composed of impure *cellulose*.

(c) Heat splinters of wood in small glass tubes, and notice the formation of *charcoal* when the air has not free access to the burning wood. Take out the charcoal, and finish the burning by heating it in a crucible.

(d) Determine the percentage of mineral matter in writing-paper.

**Starch.**—Starch is a white, powdery, non-crystallisable substance, which is formed in nearly all green plants. It is formed in the leaves and other green parts, but is afterwards stored in various other portions of the plant. It is to be found, for example, especially abundantly in seeds (notably those of wheat, oats, barley and other cereals, peas, beans, etc.) ; in the stems of such plants as the sago palm (Fig. 145) ; in what the botanist calls the *tubers* (Fig. 141) of the potato (the part which is

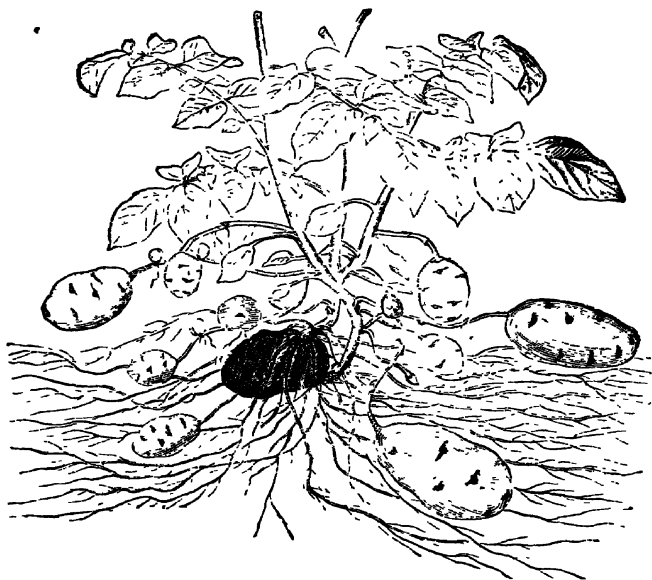


FIG. 141.—A potato plant, showing the rounded “tubers” which are used for food.

eaten) and some other plants ; and in roots. When examined microscopically, starch is seen to be composed of minute grains, which differ remarkably in size and shape, according to the plants in which they occur.

**Starch grains.**—A starch grain consists of layers arranged concentrically around a point called the *hilum*. The outer layer is the first-formed part. The largest starch grains of the potato may be  $\frac{1}{150}$  of an inch in diameter, while, on the other hand, the starch grains in beet-root are only  $\frac{1}{3200}$  of an inch in

diameter. The shape of starch grains, which is characteristic of the different plants, also varies greatly, as may be seen from Figs. 137 to 139.

**The action of water upon starch.**—Cold water has no action upon starch, but with hot water the grains swell up and burst. Their contents largely dissolve, forming a solution—*starch paste*—which when strong sets on cooling to a stiff jelly, which is used in stiffening linen as well as for other purposes.

**Starch paste is unable to diffuse through animal membranes.**—Starch paste differs from many other sub-

stances, such as a solution of sugar, in being unable to pass through an animal membrane or a sheet of wet parchment paper. This property has a very important bearing upon the digestion of starch, for no substance can be utilised by the body as food unless it is in a form which can diffuse through the walls of the digestive canal. Starch is changed into

sugar in the body before it is absorbed by the blood circulating in the walls of the lower part of the digestive canal.

**The conversion of starch into sugar.**—Under certain conditions starch becomes changed into sugars of various kinds. When starch paste is heated to boiling with very dilute acids—sulphuric or hydrochloric, for example—it undergoes a series of changes, and is at last wholly converted into a variety of sugar called *glucose*, which has the power of giving a red precipitate when heated with Fehling's solution (p. 214). A somewhat similar change takes place when starch paste is treated with malt-extract. In the latter case the transformation is brought about by a substance called *diastase*, which is produced when barley seeds germinate. The gradual corrosion of the starch grain, as



FIG. 142 - Different stages of corrosion shown by the starch grains of germinating barley. (Very highly magnified.)



it is converted into soluble sugar, is seen in Fig. 142. The sugar which is formed has, like glucose, the power of giving a red precipitate when heated with Fehling's solution, but has a different constitution and is known as *maltose*.

The saliva and the pancreatic juice of the animal body contain minute quantities of substances which act upon cooked starch in the same manner, changing it into sugar—a valuable food—and hence enabling it to pass through the walls of the digestive canal and enter the blood.

Young children, whose digestive organs are weak, are often supplied with *malted foods*.

The diastase present in these assists the action of the saliva and pancreatic juice.

**The action of iodine upon starch.**—When a solution of iodine is added to starch, a beautiful blue colour is obtained. The colour is destroyed by heating, but reappears when the liquid is cooled. The formation of this colour upon the addition of iodine affords a ready means of detecting the presence of starch.

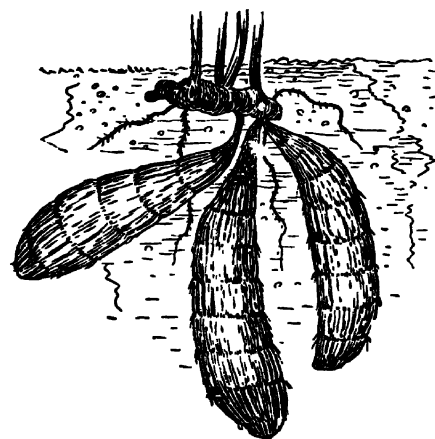


FIG. 143.—Rhizome, or underground stem of *Maranta*, from which arrowroot is prepared.

**Starch as a food.**—Starch, prepared in various ways, is extensively used as a food. Familiar examples of foods containing starch are potatoes (Fig. 141)—which consist very largely of starch—and bread. Macaroni, arrowroot, sago, semolina, rice, corn-flour, and tapioca—all popular and valuable articles of food—consist almost wholly of starch.

**Wheat starch.**—In most countries bread is the chief article of diet. To what circumstances bread owes its great value will be considered in a later chapter. It may be mentioned here that a sack of wheat flour weighing 280 lbs. contains no less than 220 lbs. of dry starch—a fact which indicates the great importance of starch as a food.

**Arrowroot.**—The purest, and also one of the most palatable and easily digested, of starches is arrowroot. It is prepared from the underground stems (or rhizomes) of a plant named *Maranta* (Fig. 143), found in the West and East Indies. The starch of arrowroot swells up very readily in hot water.

**Tapioca** is obtained from the roots of the *Cassava* (Fig. 144). The moist starch is heated with constant stirring on hot plates. During this process many of the starch grains burst and stick together, to form the small irregular masses familiar to cooks. Owing to its previous heating, tapioca is partly soluble in cold water, and readily forms a jelly with boiling water.

**Sago.**—Sago also undergoes a process of preparation before it is put upon the market. It is prepared from the pith of certain palms (Fig. 145) found in the Malay Archipelago, the Philippine Islands, and parts of India. Owing to the method of its preparation, the little balls swell up without losing their form when cooked.

An article called "Indian Sago" is manufactured in France and Germany from potato-starch. Sago-starch is sometimes used to adulterate cocoa and chocolate.

**Cooking.**—It has been seen (p. 212) that raw starch is insoluble in water. In this form, therefore, it is not easily acted upon by those digestive juices which change starch into sugar. When starch is cooked, however, the grains are ruptured, and their contents are converted into a soluble form which is readily changed into diffusible sugar. Hence all starches require to be cooked before they can be used as food.

**The use of starch for laundry purposes.**—Owing to its property of swelling up to form a stiff paste with boiling water, starch is much used for stiffening linen. In the ironing process the fabric acquires a glossy surface, which not only improves its appearance but also renders it much less liable to catch particles of dirt and become soiled. The best laundry starch is

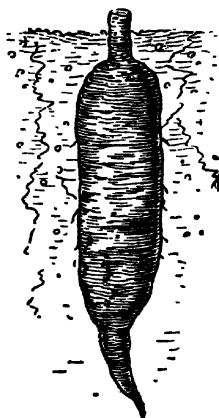


FIG. 144.—Root of the *Cassava*, from which tapioca is prepared.

made from rice, as this gives a higher lustre than any other ; but arrowroot and wheat starch are also used.

**The use of starch for manufacturing purposes.**—Starch is extensively used in the sizing and dressing of cotton and



FIG. 145.—The sago palm.

other cloths. When converted into sugar, it is largely employed in brewing, for confectionery, and in the manufacture of syrups, etc.

**Dextrin**, or "British gum," which is used in calico-printing, in the preparation of some medicines, and for gumming labels, envelopes, stamps, etc., is formed when dry starch is heated to a temperature of about 150° C. It is soluble in water, and has no action upon Fehling's solution. It gives a red colour with iodine.

**The manufacture of starch.**—The starch which is used for other than edible purposes is chiefly manufactured from wheat and rice, although potatoes are largely used in France and Germany, and maize in the United States. The grain and potatoes are ground up by machinery, and a constant stream of water washes away the starch as a milky fluid. This is allowed to settle, and the starch, after being treated with weak alkali to remove various impurities, is washed with water and finally dried.

**Starch a carbohydrate.**—Starch has the following composition :—

Carbon . . . . .	72
Hydrogen . . . . .	10
Oxygen . . . . .	80
	—
	162

Organic bodies, such as starch, which consist of carbon, hydrogen, and oxygen, and contain the last two elements in the proportion in which they occur in water (1 part of hydrogen to 8 parts of oxygen by weight) are called *carbohydrates*. Many other carbohydrates, as well as starch, are of great importance as foods.

**Cellulose** is a carbohydrate which occurs in all plants. Cellulose and starch consist of the same elements united in the same proportions. Cotton-wool and the best filter-paper are almost pure cellulose. Wood is chiefly composed of cellulose, but it contains other compounds as well. Some of these are left as ash or mineral matter when the wood is burnt.

#### SUMMARY.

**Wheat flour** contains *about* 12 per cent. of *gluten*. Most of what remains is *starch*. The gluten may be separated from the starch by kneading dough contained in a muslin bag in water.

**Gluten**, when freshly prepared, is tenacious and elastic. When dried, it becomes hard and brittle.

Gluten is a *proteid*, and is a valuable food.

**Starch** is chiefly formed in the leaves of plants, but is stored in other parts. It consists of very small *grains*, which vary in size and shape according to the plant from which the starch is obtained.

Starch is insoluble in cold water, but swells up with boiling water, the grains bursting and their contents dissolving largely to form *starch paste*.

Starch paste is unable to diffuse through moist animal membranes.

Starch may be changed into sugar :—

- i. By boiling with dilute acids.
- ii. By the action of the *diastase* of plants.
- iii. By the action of the saliva and pancreatic juice of animals.

A sugar solution passes easily through moist animal membranes.

Starch is a *carbohydrate*, and is a valuable food. It is the chief ingredient of flour, meal, rice, arrowroot, tapioca, sago, potatoes, etc. Starch is cooked before being eaten in order to burst the grains and make it soluble.

Starch is used for stiffening linen and for various manufacturing purposes.

**Dextrin**, or British gum, is formed when dry starch is heated to 150° C. It is soluble in water.

#### EXERCISES ON CHAPTER XVI.

1. What two substances constitute the food materials in flour? State the difference in their composition, and describe a process by which they may be separated. (1893.)

2. Dilute solution of iodine is added to a thin mucilage of starch, the liquid is then heated, and afterwards allowed to cool. Describe exactly what you would observe. (1898.)

3. From what sources is starch obtained? Describe the physical characters of starch, and the action of cold and hot water upon it. Describe also the action of hydrochloric acid and malt extract upon starch. (1897.)

4. What is arrowroot? Explain its value as an article of food. (1896.)

5. Describe the effects of heat upon gluten. What is the value of gluten as a food?

6. What is the action of saliva upon starch? What changes must starch undergo before it can be used by the body as a food?

7. What is dextrin? How may it be obtained from starch?

8. What is meant by the term *carbohydrate*?

## CHAPTER XVII

### THE SUGARS

#### VARIETIES OF SUGAR

**222. The properties of cane sugar.**—*Sugar is crystalline.*—Examine a lump of loaf sugar. This is the most familiar form of cane sugar. It is composed of a large number of small colourless crystals. The crystals scatter the light in all directions and the lump of sugar is therefore white. Ordinary table-salt and snow—which are also composed of small colourless crystals—are white for the same reason.

*Taste.*—Place a small piece of the sugar on the tongue and observe its sweet taste. The fact that the sugar has a distinct taste shows that it is soluble in the saliva, as it is only after substances are dissolved that their taste can be distinguished.

*Solubility.*—Find out to what extent cane sugar is soluble in cold water by putting a weighed quantity into a dish and adding water in *small* quantities at a time from a burette (Fig. 135); stir continuously. About 20 grams is a convenient quantity of sugar to take. When all the sugar has dissolved, notice how much water you have added. Remembering (p. 30) that one cubic centimetre of water weighs one gram,<sup>1</sup> it will be found that cane sugar is soluble in about half its weight of cold water.

Much more sugar can be dissolved in hot water than in cold.

*Sugar is neutral.*—Test the solution with red and blue litmus papers. There is no change in colour. Sugar is therefore *neutral*.

**223. Preparation of sugar candy.**—Make about a teacupful of a hot

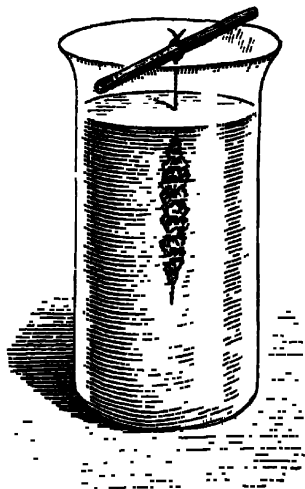


FIG. 146 —How to make sugar candy.

<sup>1</sup> See note as to temperature on p. 31.

concentrated solution of cane sugar; across the vessel containing the solution place a match or a penholder so as to suspend a piece of thin string in the solution, as shown in Fig. 146. Put the vessel aside for some hours. The sugar crystallises upon the string.

**224. The effects of heat upon cane sugar.**—Put a lump of sugar into a porcelain crucible and heat it by means of a small Bunsen flame (Fig. 87). At a temperature of about  $160^{\circ}\text{C}$ .,—or lower if a small quantity of water is present—the sugar melts to a clear liquid. Take away the flame, and allow the crucible to cool.

The sugar solidifies to a transparent brittle mass in which no crystals can be distinguished. In this condition it is known as *barley sugar*.

Take this out of the crucible, and break a piece off and taste it.

Put back into the crucible what is left, and continue the heating. At a temperature of from  $210^{\circ}\text{C}$ . to  $220^{\circ}\text{C}$ . the sugar becomes dark-brown in colour, and is then known as *caramel* or *burnt sugar*.

Take out a little of this caramel and dissolve it in water. Preserve the dark solution for Expt. 226.

**225. The presence of carbon in sugar.**—(a) Again heat the crucible and notice that the caramel blackens, catches fire, and swells up to form charcoal (*carbon*). On continued heating, this burns completely away, combining with the oxygen of the air to form carbon dioxide gas.

(b) Put about a teaspoonful of a strong solution of sugar into a large beaker, and add an equal volume of strong sulphuric acid. Stir with a glass rod. The mixture turns black and froths up greatly, giving off much steam and sulphurous fumes.

The black spongy mass consists of charcoal or carbon.

**226. How to remove the colour from burnt sugar.**—To the solution of burnt sugar made in Expt. 224, add a teaspoonful of animal charcoal. Stir for some time, until the charcoal is thoroughly mixed with the liquid, and then filter through filter-paper in a funnel (Fig. 53).

The action of the charcoal has made the liquid, which runs through the filter, either colourless, or at least very much paler in colour than the caramel solution which was taken.

**227. The behaviour of cane sugar with Fehling's solution.**—To about a teaspoonful of Fehling's solution in a test-tube, add a crystal of cane sugar, and warm as in Expt. 215.

The Fehling's solution remains blue, no red precipitate being produced with cane sugar.

**228. The properties of glucose.**—Examine a specimen of commercial glucose. It is a thick, colourless syrup, which mixes with water in all proportions.

Taste a small quantity. It is sweet, but not nearly so sweet as cane sugar.

**229. The behaviour of glucose with Fehling's solution.**—Heat a little glucose with Fehling's solution. The blue colour of the solution disappears almost at once, and the liquid becomes turbid with a red precipitate.

We have already seen (Expt. 216) that glucose can be prepared

from starch by heating the latter substance with dilute sulphuric acid.

**230. The properties of grape sugar.**—Obtain a quantity of “granulated grape sugar.” It consists of crystals, and looks very much like ordinary granulated cane sugar. Like it, grape sugar is soluble in water, though less freely.

Taste a small crystal of each, and notice that grape sugar is less sweet than cane sugar.

**231. The behaviour of grape sugar with Fehling’s solution.**—Add a small crystal of grape sugar to about a teaspoonful of Fehling’s solution in a test-tube, and warm. A red precipitate is obtained.

**232. Grape sugar is present in honey.**—Put a little honey into the mouth, and compare its sweetness with that of cane sugar and grape sugar respectively.

Also notice that the honey contains solid particles which crush easily between the teeth. Separate the solid particles from the syrupy part of the honey by stirring up some honey in a dish or beaker with spirit, and pouring off the spirit. Small crystals are left. Test the crystals with Fehling’s solution in the usual way. A red precipitate is produced.

**233 The inversion of cane sugar.**—Dissolve in a teacupful of water as much sugar as will lie on a shilling. We know from Expt. 227 that this solution will have no action upon Fehling’s solution warmed with it. Put the sugar solution into a flask, and add to it about a

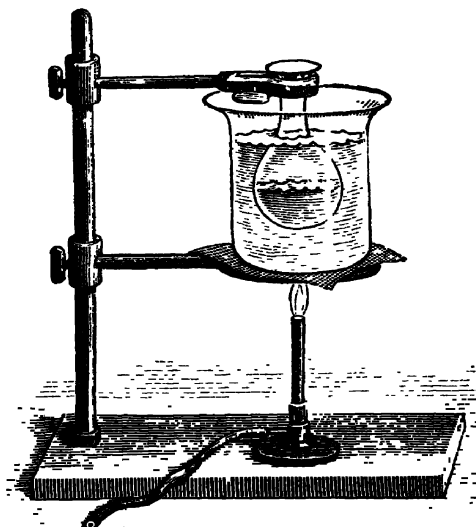


Fig. 147.—Cane sugar is “inverted” or changed into glucose by prolonged heating with an acid.

dozen drops of strong hydrochloric acid. Place the flask in boiling water, as shown in Fig. 147, and keep the water boiling for about half an hour. At the end of this time put the contents of the flask into a dish, and add sodium carbonate in small quantities at a time until the liquid is just neutral, that is, until a drop taken out on the end of a glass rod, after stirring, just fails to redden blue litmus paper, and yet does not turn red litmus paper blue. Test a little of this neutral solution with Fehling’s solution. A red precipitate is produced.

During the experiment the cane sugar has been “inverted,” that is, changed into glucose. Glucose which has been prepared by the inversion of cane sugar is called *invert sugar*.



**Cane sugar.**—Cane sugar is in common use in all civilised countries for sweetening food. It derives its name from the plant—known as the sugar cane (Fig. 148)—from which it was first obtained. The knowledge of this plant was first brought to Europe from Syria by the Crusaders. Before that time honey was universally used for sweetening food, but the sweet juice of

the sugar cane had a flavour which rapidly gained favour among those who were able to afford the new luxury.

Soon after the beginning of the fifteenth century, it became usual to evaporate the juice of the sugar cane in order to obtain the solid sugar; and a hundred years later manufacturers introduced a process of refining, by which the appearance and quality of the sugar were greatly improved.

Until the middle of the eighteenth century all the cane sugar used was obtained from the sugar cane, which was cultivated chiefly in Central America and the West Indies. In 1747, however, Marggraf discovered that cane sugar occurs also in beetroot (Fig. 149), and since that time the prepara-



FIG. 148.—The sugar cane.

tion of sugar from beetroot has increased enormously. At the present time some four million tons of sugar are manufactured annually in Europe from beetroot. Moreover, improved methods of cultivation have largely increased the amount of sugar in beetroots, so that 100 lbs. of sugar are now obtained from 750 lbs. of beetroot.

**The manufacture of sugar.**—In the modern methods for

the extraction of sugar, the beetroots are first cleaned and then cut into slices, and the sugar dissolved out by a continuous stream of warm water. When the sugar cane is used—and one-third of the world's supply of sugar is still obtained from cane—the canes are either crushed between rollers to squeeze out the juice, or are cut and broken into small pieces, which are treated with warm water to dissolve out the sugar. Lime is added to the strong solution of sugar, and the mixture heated almost to boiling, when most of the impurities separate as a scum upon the surface. The clear liquid beneath is separated from the scum, and sent into another tank, where carbon dioxide gas is forced into it. This combines with the lime which is in solution, forming a precipitate of chalk (p. 157), which settles to the bottom, carrying with it almost all the remaining impurities.

When the juice is clear, it is run off and passed through animal charcoal. The charcoal to a very large extent retains the colouring matter, and the liquid which flows away is concentrated by evaporation. As a high temperature would prevent the sugar from afterwards crystallising properly, the evaporation is carried out in what are known as *vacuum pans*. The reduced pressure enables the evaporation to take place at a much lower temperature than would be necessary in open pans.

When the liquid is sufficiently concentrated, it is allowed to cool, and is then run into conical moulds. The sugar separates in the form of a mass of crystals, which, after being drained and dried, forms *loaf sugar*.

"Soft sugar" is obtained from the less concentrated solutions.

The liquor which remains is known as *molasses*. It still contains a large amount of sugar (about 50 per cent.), and it is only recently that it has been found possible to separate this



FIG. 149.—Beetroot.

The molasses from beet-sugar has an objectionable flavour, and is never used for table purposes.

Cane sugar occurs not only in the sugar cane and in beet-roots, but also in the Java palm, the sugar maple, carrot, turnip, and madder roots, coffee, various nuts, flowers, honey, etc.

When pure, cane sugar forms colourless crystals. It has a very characteristic sweet taste, and is easily soluble in water. Cold water dissolves about twice its own weight of sugar, and hot water dissolves a still greater proportion.

**Sugar candy.**—When cane sugar is allowed to crystallise slowly upon threads from a strong solution, sugar candy is obtained. In this form the appearance of the sugar crystals can be well observed (Fig. 150).



FIG. 150.—Crystals of sugar candy.

**Barley sugar.**—Cane sugar melts, upon heating, at a temperature of about  $160^{\circ}\text{C}$ ., or, if water is present, at a lower temperature. If this melted sugar is cast into moulds, a transparent solid known as barley sugar is formed.

**Caramel.**—Upon heating to  $210^{\circ}\text{C}$ ., the sugar darkens in colour, and is converted into a dark brown substance known as caramel or burnt sugar. A solution of caramel is much used for colouring sauces, etc. It is, for example, added to artificial vinegar to give it the brown colour which appears to be regarded as a sign of "strength" in such liquids. Burnt sugar has a very characteristic odour.

When cane sugar is still more strongly heated, it becomes black, catches fire, and much of its carbon separates as a very pure form of charcoal. On continued ignition in an open vessel it burns completely away.

**The composition of cane sugar.**—Cane sugar consists

of carbon, hydrogen, and oxygen united in the following proportions :—

Carbon . . . . .	72
Hydrogen . . . . .	11
Oxygen . . . . .	88
	<hr/>
	171

We have already learnt (p. 221) to apply the name *carbohydrates* to bodies which, like starch and cane sugar, consist of carbon, hydrogen, and oxygen : the last two elements being present in the proportion (1 part of hydrogen to 8 parts of oxygen by weight) in which they occur in water.

When carbohydrates burn, their carbon combines with oxygen from the air to form carbon dioxide gas, and their hydrogen combines with oxygen to form water.

**The saccharoses.**—The composition of cane sugar, or saccharose, just given, is also that of some other sugars. Of these the commonest are milk sugar, or *lactose*, and *maltose*.

*Lactose* occurs in milk. It is prepared by evaporating the liquid (whey) which is left when milk is curdled by means of rennet. The lactose is allowed to crystallise on strings or splinters of wood. It is hard and granular, and only slightly sweet.

*Maltose* is the sugar which is formed when diastase is allowed to act upon starch (p. 217).

The sugars which have the same composition as cane sugar are collectively known as the *saccharoses*.

**The inversion of cane sugar.**—A solution of cane sugar undergoes a remarkable change when a few drops of hydrochloric acid are added to it and the mixture is kept for half an hour at the temperature of boiling water. Before the experiment a few drops of the solution have no action upon Fehling's solution (p. 214) when warmed ; but after heating with hydrochloric acid, the solution very readily produces a red precipitate when it is warmed with Fehling's solution.

The explanation of this difference in behaviour lies in the fact that the cane sugar has been changed by the treatment into a form of sugar commonly called *glucose*. The change is

known as "inversion," and the glucose so produced is known as "invert sugar."

**Glucose or grape sugar.**—Glucose or grape sugar occurs naturally in the nectar of flowers (honey) and in sweet fruits. It may be seen in raisins, for example, as small yellow grains. As cane sugar is found in unripe fruits, it seems likely that grape sugar is produced during the ripening of fruits by a change of the nature of inversion.

Crystals of grape sugar may easily be obtained from honey by washing away its more fluid parts with spirit.

Crystallised grape sugar resembles cane sugar in appearance, but it is much less sweet: the sweetness of grape sugar having to that of cane sugar about the proportion of 2 to 5.

Glucose is prepared artificially by boiling starch with dilute sulphuric acid. After the excess of acid has been removed by adding chalk and filtering off the calcium sulphate which is produced, the solution of glucose is concentrated by evaporation.

If the evaporation is only carried to the extent necessary to produce a thick syrup, the product is known commercially as *glucose*. It is used in brewing and in the finishing of cloth, for the preparation of jam, table-syrups, etc., and for other purposes.

When glucose is still further concentrated, it forms white crystals.

A very characteristic test for glucose or grape sugar lies in its behaviour with Fehling's solution. The presence of a very small quantity of this form of sugar causes the solution, on warming, to become turbid and deposit a red precipitate. As cane sugar does not cause the separation of this precipitate from Fehling's solution, the amount of glucose present in a mixture of the two forms of sugar may be determined by collecting and weighing the precipitate.

The different forms of sugar have a very remarkable action on light of a particular kind ("polarised" light). By taking advantage of this property, it is possible to find the strengths of solutions of sugar with less trouble than by using Fehling's solution. An instrument called a *polarimeter* is used for this purpose.

**The glucoses.**—What is generally known as "glucose" really consists of a strong solution of a mixture of different sugars,

most of which, however, have the following composition :—

Carbon . . . . .	72
Hydrogen . . . . .	12
Oxygen . . . . .	96
	<hr/>
	180

This should be compared with the composition of cane sugar, given on p. 229.

The compounds having this composition (which, like all sugars, are carbohydrates) are classified together as the *glucoses*; just as the compounds of the cane sugar type are collectively known as the *saccharoses*. The sugars thus fall naturally into these two groups.

**Sugars as foods.**—By themselves the sugars are incapable of supporting life; but, as soluble and easily digested carbohydrates, they are nevertheless of very great importance as food-stuffs. This question will be more fully considered in a later chapter; but we may here call attention to the use of sugar not only as a food in the strict sense of the word, but as a sweetening agent, which increases the attractiveness of other foods, and thus assists digestion.

We are, perhaps, somewhat inclined to under-estimate the influence of emotional conditions upon digestion and health. From this point of view alone, the value of sugar is very high.

**Saccharine.**—It is convenient to mention here a complex organic compound called *saccharine*. This is not in any sense a sugar (nor has it any value as a real food), but belongs to a quite different class of compounds. It is remarkable in possessing 300 times the sweetening power of cane sugar, so that it is used in very small quantities at a time. Its great importance, from a dietetic point of view, lies in the fact that, owing to its composition, it may be used by patients suffering from diseases—notably diabetes—in which carbohydrates are distinctly injurious. The discovery of saccharine, therefore, allows such patients to enjoy sweetened dishes from which they would otherwise be debarred.

## SUMMARY.

**Cane sugar** is a white crystalline substance, which is very soluble in water, and has a sweet taste. It is chiefly obtained from beetroot and from the sugar cane. By melting and re-cooling, it is converted into a non-crystalline form called *barley sugar*.

When boiled for a long time with a little acid, cane sugar is converted into "*glucose*." Glucose so formed is called *invert sugar*. Crystallised glucose is called *grape sugar*.

**Grape sugar** is only  $\frac{2}{3}$  as sweet as cane sugar. It occurs naturally in sweet fruits, honey, etc.

The sugars are *carbohydrates*. They are not only soluble, but also pass readily through moist animal membranes, and are hence very valuable foods.

## EXERCISES ON CHAPTER XVII.

1. Describe the properties of cane sugar. What is the effect of heat upon cane-sugar?
2. Describe how you would proceed to make "sugar candy." What is the difference between "sugar candy" and "barley sugar"? (1898.)
3. How would you demonstrate the presence of carbon in sugar?
4. In what plants does cane sugar occur in nature? By what processes is it prepared?
5. How does grape sugar occur in nature? How does it differ from cane sugar?
6. How can grape sugar be made from (a) starch, (b) cane sugar?
7. What are the physical and chemical characteristics of carbohydrates? (1898, A.)
8. What advantages as a food has sugar over starch?
9. What compounds are formed when sugar is burnt in air?

## CHAPTER XVIII

### ALCOHOL AND ALCOHOLIC BEVERAGES

#### THE PROPERTIES OF ALCOHOL

**234. Rectified spirit.**—For these observations of the properties of alcohol, rectified spirits of wine may be conveniently employed. This consists of a mixture of alcohol and water, in the proportion of 84 parts of alcohol to 16 parts of water.

*Colour and mobility.*—Hold the bottle up to the light, and notice that the liquid is colourless. It flows more readily, when the bottle is slightly moved, than water would. It is, therefore, said to be more *mobile* than water.

*Smell and taste.*—The smell of alcohol is very strong and characteristic, and is readily recognised. Place a few drops on the tongue. The spirit has a burning taste.

*Solubility in water.*—Add about a teaspoonful of spirit to a cupful of water. The spirit does not form a separate layer, as oil would, but dissolves completely.

Mix equal quantities of spirit and water in a dish, and notice that in this case also a clear liquid is obtained, in which the alcohol and water cannot be distinguished from each other. The mixture is distinctly warm.

Alcohol and water are soluble in each other in all proportions, heat being liberated during the process of solution.

*Neutrality.*—By means of a glass rod place a drop of spirit on red and blue litmus papers. There is no change of colour. Alcohol is therefore neutral.

*Volatility.*—Pour a few drops of spirit on the palm of the hand, and rub it with the finger. A sensation of cold, due to the rapid evaporation of the spirit, is felt.

Put about a teaspoonful of spirit into a test-tube, dip a thermometer into it, and place the test-tube in hot water. Notice that the spirit begins to boil at a temperature considerably below that of boiling water.

*Caution!* Be careful not to set the spirit on fire.

*Inflammability.*—Put a teaspoonful of spirit into an evaporating basin, and apply a light. The spirit catches fire, and burns with a pale blue flame. Put a piece of white porcelain into the flame, and notice



that there is no deposit of soot, such as would be given by a candle or match flame (p. 179).

The spirit burns away without leaving any residue whatever.

**235. Carbon and hydrogen present in alcohol.**—Set fire to a few drops of spirit in the cup of a deflagrating spoon, and lower the spoon at once into a clean glass jar (Fig. 90). Notice the "dew" which forms upon the sides of the jar. When the flame has gone out, remove the spoon, pour a little lime-water into the jar, and shake up. The lime-water becomes milky, owing to the presence of carbon dioxide (p. 163). The carbon dioxide is formed during the burning of the spirit, by the combination of its carbon with the oxygen of the air, and the water of the "dew" which dims the sides of the jar is formed by the oxidation of its hydrogen.

**236. The solvent power of alcohol.**—(a) Put a teaspoonful of castor oil into a dish, and add five or six times its volume of spirit. The oil

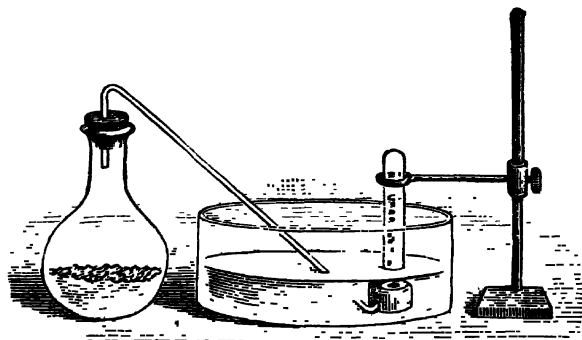


FIG. 151.—When a warm solution of sugar is fermented by yeast, alcohol and carbon dioxide are formed.

dissolves completely in the spirit, forming a perfectly uniform liquid. Use water instead of spirit, and notice that the oil remains separate as a layer floating on the water.

(b) Try whether resin and sugar dissolve in spirit. Compare the solubility of these solids in spirit with their behaviour with water.

**237. The specific gravity of mixtures of alcohol and water.**—By the method explained on p. 32 find the density, or specific gravity, of spirits of wine. It is about 0.84. [What is the weight of a litre of a liquid having this specific gravity?]

Find also the specific gravity of a mixture of (a) about two parts of spirit to one of water, (b) about one part of spirit to two of water.

The specific gravity increases with the proportion of water present.

**238. The formation of alcohol from glucose or grape sugar.**—(a) Dissolve about half an ounce of grape sugar in a quart of water. Put the solution into a large flask, fitted up with a cork and delivery tube as shown in Fig. 151, and add a little brewer's yeast. Keep the flask in a warm place for some hours. The mixture begins to froth, and

seems to boil owing to the large quantity of gas which is given off. When the air has been expelled from the apparatus, some of the gas may be collected in a test-tube over the pneumatic trough, as little water as possible being used.

When the tube is full of gas, remove it and test the gas. It will be found to extinguish a lighted taper placed in it, and to turn lime-water milky. The gas given off has no smell, and is therefore carbon dioxide.

(b) Notice also that the liquid in the flask has become very turbid during the experiment, the quantity of yeast having increased.

(c) Connect the flask with a condenser (Fig. 61), and distil over about one-third of the contents. Notice the smell of the liquid which distils over. It is not so pleasant as the smell of spirits of wine, but it is nevertheless distinctly alcoholic.

**239. The yeast plant.**—Take a little speck of the kind of yeast used in Expt. 238, and mix it with a drop of water on a microscope-slide. Cover the drop with a cover-glass, and examine it with the low and high powers of the microscope. Notice the oval bodies of which the yeast is composed (Fig. 153). Each of these oval bodies is a living plant, capable—when placed in a dilute solution of sugar—of growing and multiplying, and changing some of the sugar into alcohol and carbon dioxide gas.

**240. Wild yeast.**—Make about a cupful of a dilute solution of grape sugar, and set it aside for a day or two, leaving it freely exposed to the air. Then put it in a warm place, and notice that it soon begins to “work.”

The “working,” or *fermentation*, of the liquid is due to yeasts which have fallen into it in the form of dust. These small plants are very widely distributed in the air, and readily gain access to food, etc., which is left uncovered.

Yeast plants which occur in dust are called *wild yeasts*, to distinguish them from species which are specially cultivated for their useful properties.

**241. Alcohol in beer.**—Pour a quart of beer backwards and forwards for several times from one vessel to another. This causes the escape of much of the carbon dioxide gas which is dissolved in the beer, and lessens the tendency of the liquid to froth up when heated. Put the beer into a large flask, and distil off about half of it. Notice the smell of the liquid which collects in the receiver. It is a mixture of alcohol and water. To remove the water from the liquid, add a handful of quicklime. Cork the flask and shake. After allowing it to stand for some hours, distil again; this time heat the liquid by means of boiling water contained in a vessel surrounding the distilling-flask. The quicklime retains most of the water, so that the liquid which now distils over is fairly strong alcohol.

Test it by taste and smell, and see if a little placed in a dish catches fire on the application of a light.

**Alcohol.**—From very early times, and in all parts of the world, the human race has been addicted to the use of

fermented beverages. In spite of the great differences, in methods of preparation and other respects, of these various liquids, they all owe their attraction to one and the same active principle—alcohol.

**Properties of alcohol.**—Alcohol is a neutral, colourless liquid, possessing a peculiar and—to most people—pleasant odour. It has a burning taste. It is soluble in all proportions in water, heat being liberated when the liquids are mixed. Taken in dilute solutions alcohol produces intoxication, but is poisonous when drunk pure or in strong solution.

Alcohol is readily inflammable, burning with a pale blue flame which yields no soot. It is therefore a convenient fuel

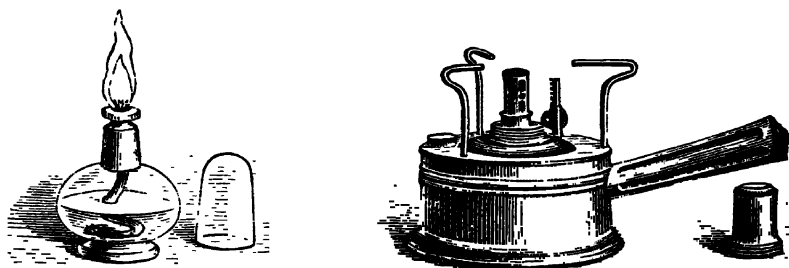


FIG. 152.—Two kinds of spirit lamp.

for heating vessels in cases where a smoky flame is undesirable, or where coal gas or oil cannot be obtained. Several *spirit lamps* (Fig. 152) of various forms have been devised for the burning of alcohol.

Pure alcohol is considerably lighter than water, having a specific gravity of about 0.8. It dissolves many organic substances which are insoluble in water.

**The separation of alcohol and water.**—Alcohol boils at a temperature of  $78^{\circ}5$  C. When, therefore, a mixture of alcohol and water is distilled, the liquid begins to boil at a temperature intermediate between this and  $100^{\circ}$  C., the boiling point of water. The first portions evaporated contain a greater proportion of alcohol than the last liquid distilled over, and by repeated distillation the strength of the liquid can be increased so as to obtain at last an almost complete separation of the

alcohol from the water. The separation is more thorough if in the later stages the liquid is allowed to stand over quicklime for some hours before being distilled. The quicklime combines with the water, and retains most of it, so that, upon heating, the liquid which is driven off and condensed is almost all alcohol.

**The composition of alcohol.**—Pure alcohol has the following percentage composition :—

Carbon . . . . .	52.17
Oxygen . . . . .	34.78
Hydrogen . . . . .	13.04
	<hr/>
	99.99

The presence of carbon and hydrogen in the compound may be readily shown by burning alcohol in a deflagrating spoon in a dry glass jar. During the burning of the alcohol, its hydrogen is oxidised to water (p. 143), and its carbon to carbon dioxide gas (p. 127). The water may be seen as a slight dew upon the sides of the jar. The presence of the carbon dioxide may be shown by the milkeness produced when a little clear lime-water is shaken up in the jar.

**The formation of alcohol.**—Although chemists are acquainted with methods by which alcohol can be directly built up, as it were, from the elements of which it is composed, it is in practice always produced by the fermentation of one or other of the various sugars.

This may be well illustrated by making a dilute solution of grape sugar in water, adding a little yeast, and keeping the mixture in a warm place. It soon begins to “work,” and becomes frothy and turbid. Three chief results of the action may be readily observed :—

(a) The quantity of the yeast increases.

(b) A gas is given off, which may be collected, tested, and proved to consist of carbon dioxide.

(c) Alcohol is formed in the liquid.

**Yeast.**—The cause of this remarkable change of sugar into alcohol and carbon dioxide is found in the yeast which is added. When examined by means of the microscope, yeast is seen to be composed of a mass of tiny oval bodies called *cells* (Fig. 153). Each of these is a living plant. The yeast plant

commonly reproduces its kind by a process of "budding," in which a knob forms on the cell, increases in size, and is at last cut off from the parent to become a separate cell. The cells often remain in contact for some time before finally breaking off.

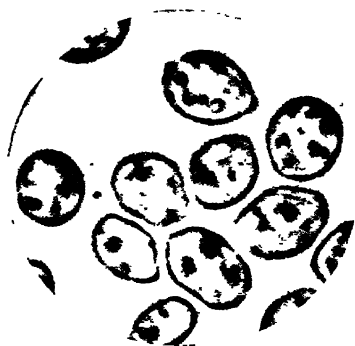


FIG. 153.—Cells of yeast. (Highly magnified.)

**Ferments and the nature of Fermentation.**—When a sugar solution is "worked" with yeast, the evolution of gas often so great as to give an appearance of boiling. The name *fermentation* was consequently applied to the changes taking place. The term is still used for all processes of a similar nature, even when there is no disengagement of gas, and

therefore no appearance of boiling, such as occurs in alcoholic fermentation.

For example, when beer and wine turn sour, and become converted into vinegar by the formation of acetic acid, the change is brought about by the activities, not of yeast, but of another microscopic plant—the *vinegar organism*, shown in Fig. 158. Here the change is known as *acetic fermentation*.

Another kind of fermentation is exemplified by the action of the diastase of malt-extract (p. 217) in converting starch into sugar. Here the action is not due to living cells, but to a complex organic compound (*diastase*) formed by the seed.

As the yeast plant, the vinegar organism, and diastase bring about fermentation, they are called *ferments*. They resemble each other in one very important respect: they cause the chemical change of an amount of material which is out of all proportion to their own quantity. In alcoholic fermentation, for example, while the yeast-cells use up a small quantity of sugar for their own nutrition, they cause an enormously greater amount of sugar to be changed into alcohol. Again, the diastase of malt-extract does not itself appear to undergo any change, notwithstanding the extensive conversion of starch into sugar which it induces.

It is a curious fact that the activity of ferments stops when the substances they produce reach a certain degree of concentration.

**Organised and unorganised ferments.**—It would seem, from what has been said in the previous paragraph, that ferments are of two kinds : firstly, living cells, such as yeast and the vinegar organism ; and secondly, organic compounds, which, like diastase, are not themselves alive, but are produced by living plants or animals. It is usual to express this distinction by terming ferments of the first class *organised*, and those of the second class *unorganised ferments* or *enzymes*.

The belief is, however, gaining ground that even organised ferments really act by virtue of enzymes which their living cells produce. The evidence in favour of this view is very strong, but cannot be considered here.

**Beer and brewing.**—Barley contains about 55 per cent. of starch, and this starch is made use of in the process of brewing. The barley is first moistened, and then kept warm until it begins to sprout. In these early stages of growth a small quantity of the ferment diastase is produced. The growth of the grain is stopped at the proper stage by heating it in kilns—a treatment which converts the barley into *malt*. The malt is then crushed and treated with water for about two hours at a temperature of from 50° to 70° C. During this *mashing* process the diastase acts upon the starch (Fig. 142), and largely converts it into the variety of sugar known as maltose (p. 229). Hops are then added, so as to give the beer a bitter taste, and the liquor or *wort* is boiled and strained.

After the wort has been rapidly cooled, the necessary quantity of yeast is added and alcoholic fermentation is allowed to take place. When this has proceeded far enough, the yeast is removed from the beer, and the latter is stored in casks or bottles. The trace of yeast which is still present is sufficient to cause a slow additional fermentation, so that the carbon dioxide—unable to escape—gradually accumulates in the stored beer.

The dark colour of *porter* or *stout* is due to the use of malt which has been charred in the heating.

Barley is most commonly used for beer-making, but it is evident that a fermented beverage of the nature of beer can be similarly prepared from any grain rich in starch. The Berlin

*Weissbier* is made from wheat ; and maize, millet, rye, and rice are also used in different parts of the world for similar purposes.

It should be noticed that the process of beer-making outlined above consists essentially of two fermentations\*. (*a*) the *diastatic* fermentation, by which starch is changed into sugar, and (*b*) the *alcoholic* fermentation, whereby the sugar is converted into alcohol and carbon dioxide.

To save the time and trouble involved in malting and mashing, it has become a common practice with many brewers to use commercial glucose. This substance is manufactured by the action of sulphuric acid upon starch (p. 230), and unless purified acid is employed, the glucose is naturally liable to contamination. The disastrous effects of using for this purpose crude sulphuric acid, containing arsenic, were seen recently (July 1900 to January 1901), when a large number of persons were poisoned by drinking beer brewed from glucose.

*Mead*, a drink which was formerly very common in Northern Europe, is made by fermenting a solution of honey.

The use of *Koumiss*, or fermented mares' or asses' milk, is practically confined to Central Asia. Milk sugar, or lactose (p. 229), like cane sugar, is not directly fermentable ; it is, however, changed into a fermentable form of sugar by the action of the lactic acid which is formed in the milk.

Similarly, it has been found that the action of yeast upon cane sugar is first to invert it—that is, to change it into grape sugar. The grape sugar thus produced then ferments, forming alcohol and carbon dioxide.

**Wines.**—Grape juice very easily undergoes alcoholic fermentation, for it not only contains glucose or grape sugar, which is directly fermentable, but an alcoholic ferment occurs on the skins of the grapes themselves, rendering the addition of yeast unnecessary. The liquid which is obtained as a final result of the fermentation of grape juice is known as *wine*.

Wines vary greatly in the amount of alcohol and other substances they contain, in colour, and in the aroma or *bouquet*. These differences are due, not only to the kinds of grapes used, but also to the methods of manufacture and other circumstances.

**Spirits.**—The commonest spirits are brandy, rum, whisky, and gin. *Brandy* is made by distilling the spirit from certain

*white* wines. *Rum* is the spirit distilled from fermented molasses (p. 227). Brandy and rum owe their special flavours to certain complex compounds which distil over with the alcohol. *Whisky* is generally made by the distillation of fermented malt-liquors, although rye is sometimes added to the malt. *Gin*, a spirit made from malt and rye, contains various added flavouring substances.

**Amount of alcohol in various beverages.**—The average percentage quantity of alcohol in some of the commonest liquors is shown in the following table:—

Whisky . . . . .	60 per cent.
Brandy . . . . .	50    "
Port . . . . .	18    "
Champagne . . . . .	11    "
Claret . . . . .	9    "
Ale . . . . .	6    "
Lager beer . . . . .	4    "

**The manufacture of alcohol.**—The first stages in the manufacture of alcohol generally consist in malting barley and mashing with water in the manner already described for the brewing of beer. After fermentation, the liquor is pumped into large *stills*, and repeatedly distilled until the required strength is obtained. A large quantity of spirit is now also manufactured from potato starch. The starch is first converted into glucose, and this is then fermented.

**Recognised strengths of alcohol.**—Special names are given to mixtures, in certain proportions, of alcohol and water. *Absolute alcohol* is supposed to be free from water. *Rectified spirits of wine* contains 84 per cent. of alcohol. *Proof spirit* contains 49·24 per cent. of alcohol. The term *proof* is derived from the old method which Excise-officers had of testing spirit by moistening gunpowder with it, and then applying a light. Proof spirit was the weakest mixture which would ignite the gunpowder when the alcohol was all burnt off. *Methylated spirit* is simply rectified spirit which has been mixed with wood spirit and mineral naphtha to make it undrinkable. It is sold duty-free, and is largely used for heating purposes and in the arts.

**The use of alcohol in the arts.**—Alcohol, chiefly owing



to its great solvent power and volatility, is much used for the preparation of varnishes, and for tinctures, perfumes, etc.

**The value of alcohol as a food.**—To people in ordinary states of health, alcohol has practically no value as a food ; and its regular use has distinctly injurious effects upon the digestive and nervous systems. On the other hand, as a stimulant in cases of illness and collapse, it is often of the greatest value ; and, owing to the readiness with which it is oxidised (p. 233), may serve as a source of energy to patients who are too weak to digest ordinary food.

#### SUMMARY.

**Alcohol** is a colourless, neutral, mobile liquid, with a pleasant smell and a burning taste. It is composed of carbon, hydrogen, and oxygen. It is very *inflammable*, and burns with a smokeless flame. It dissolves many substances which are insoluble in water.

Alcohol mixes with water in all proportions. It is very *volatile*, and boils at  $78^{\circ}5$  C. It can therefore be separated from a mixture of alcohol and water by repeated distillation.

The density of pure (or "absolute") alcohol is only  $\frac{4}{5}$  that of water.

Alcohol is the intoxicating principle in beers, wines, and spirits. It is always produced, together with carbon dioxide, by the *fermentation* of grape-sugar by yeast. Cane-sugar is not fermented until it has been first converted by the yeast into grape sugar.

#### EXERCISES ON CHAPTER XVIII.

1. How would you obtain alcohol from beer? By which of its chemical and physical properties would you distinguish alcohol from water? (1899).

2. By what processes may starch be rendered soluble in water? What becomes of the starch of barley in making beer? (1899.)

3. Describe what happens when yeast is added to a solution of cane sugar. What is the first change? How may alcohol be obtained from the product? (1899.)

4. Describe the process of brewing, and the theory of fermentation. (1894.)

5. What is a fermented drink? Explain its general preparation and effects. (1897.)

6. What is the composition of "spirits of wine"? How would you show that it boils at a lower temperature than water? What compounds are formed when the spirit burns in air? What is "proof spirit," and why is it so called? (1895.)

7. Explain what is meant by the terms (1) solvent action ; (2) distillation ; (3) chemical combination ; (4) neutralisation ; (5) fermentation. (1894.)

8. What two ferments are especially concerned in the brewing of beer ; and what chemical changes do they bring about?

## CHAPTER XIX

### A LOAF OF BREAD

#### MANUFACTURE AND PROPERTIES OF BREAD

**242. Starch and gluten.**—Before commencing the study of this chapter, the student should again separate a little flour into its chief constituents, starch and gluten, as described in Expts. 197 and 198.

**243. How to make a loaf.**—Take half a pound of flour in a basin, and add a saltspoonful of salt. Make  $\frac{1}{2}$  oz. of yeast into a thin cream with warm water, adding about half a teaspoonful of ordinary sugar. Make a hollow in the middle of the flour, put in the yeast, adding as much more tepid water as is sufficient to make a light dough on kneading the mixture. After kneading, cover the basin with a cloth, and put in a warm place for an hour. The dough swells or “rises” into a kind of sponge, which occupies at least twice its original space. Cut into pieces (putting one aside for Expt. 245), and make into “rolls,” which should be placed on a greased iron plate, and again allowed to “rise” for at least fifteen minutes. Put the plate into a hot stove (or an air-oven may be made to answer) until the rolls are baked.

**244. The appearance of new bread.**—Allow the rolls to cool, and then break one across and examine it. The inside is composed of white spongy crumb, while the crust is hard and crisp, and darker in colour.

**245. Sour dough or leaven.**—Put aside for a few days the piece of dough saved from Expt. 243 until it becomes sour, and then try the effect of using it *instead of yeast* for mixing with another quantity of flour. Make a few rolls, and compare their taste and appearance with those made in the ordinary manner.

**246. The use of baking powder.**—Make another batch of rolls in the same way as before, but use about half a teaspoonful of baking powder instead of the yeast and sugar. Compare the resulting bread with that made in Expt. 243.

**247. Unleavened bread.**—Make cakes, using neither yeast, baking powder, nor sour dough (leaven), and notice that they do not “rise” on being kept in a warm place. Bake them, and compare their taste and appearance with those of ordinary bread.

**248. The action of baking powder.**—Put a little baking powder into a test-tube, and add water. Notice that a gas is given off, causing effervescence. Test the gas by means of lime-water, as in Expt. 169.

The lime-water turns milky, showing that the gas is carbon dioxide.

In Expt. 238 it was found that carbon dioxide was also given off when yeast acted upon a warm dilute solution of sugar. It seems likely therefore that the "rising" of bread is due to the evolution of carbon dioxide by the yeast or baking powder, whichever is used.

**249. Bread-making from the beginning.**—Take a small handful of wheat-grains, and grind the grains as finely as possible in a mortar. Make the flour obtained into a dough with water, and bake in the



FIG. 154.—The primitive method of grinding corn.

oven. Taste the bread. This is the most primitive method of making bread.

**250. Mineral matter in bread.**—Weigh a porcelain crucible, first empty, and then half full of bread crumbs. The difference between the weights is the amount of bread taken. Carefully ignite the crucible over the Bunsen burner (Fig. 87), until nothing is left in it but a little white ash. Allow the crucible to cool, and weigh again.

From the weight of ash obtained calculate the percentage of mineral matter in the bread taken.

The mineral matter of white bread amounts to about 1 per cent. It is largely composed of phosphates.

**Bread.**—It is universally recognised that, next to milk, our most generally useful food is furnished by the seeds of the various cereals. Fortunately, these grasses may be grown with little trouble, and in most parts of the world. We consequently find that bread, prepared from the grain of wheat (Fig. 155), rye, barley, oats, and other cereals, has formed one of the staple foods of mankind during the whole period of which there is any historical record; and no other food has yet been discovered which is at the same time so palatable, so nutritious, and so easily obtained. There have, of course, been great improvements in the methods of bread-making, but the principle involved is essentially the same, whether the grain is reduced to flour by a complicated milling process, or ground between two flat stones (Fig. 154), in the manner described in the Bible; or whether the final product is a dainty French roll, or the sodden "damper" which the squatter bakes in the ashes of his camp-fire.

The grain is first ground to powder, and the coarser parts, or *bran*, are sifted from the fine flour. In bread-making, the flour is kneaded into a paste with water, a certain amount of yeast and sugar is usually added, and the dough is set in a warm place to "rise." Bubbles of carbon dioxide gas are given off by the activity of the yeast, and the dough becomes spongy. The mass is then heated to a much higher temperature, which kills the yeast, expands the bubbles of gas, and at last forms the bread, with its crisp outer crust, and firm, spongy interior.

**Flour.**—Fig. 156 is a diagram of a thin slice taken along the length of a grain of wheat. The grain is oval in shape, and is divided by a flat plate, the *scutellum*, into two unequal parts. The upper part is filled with a substance known as *endosperm*; while the lower portion, placed somewhat to one side of the grain, consists of the *embryo*—destined, under favourable conditions, to give rise to a new wheat plant. At the upper end of



FIG. 155.—Wheat.

the embryo the young stem and at the lower end the young root are already to be distinguished, though they are still enclosed in protecting sheaths. The whole grain is surrounded by six layers (not all shown separately) which together constitute the *bran*. The endosperm consists of a store of food formed for the support of the embryo during the earlier stages of its growth. The plate—the scutellum—which lies between

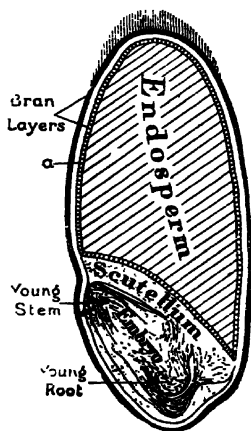


FIG. 156.—Diagram of a grain of wheat cut through from end to end. The endosperm is made into flour.

the embryo and its food, acts as a nurse to the tiny plant—absorbing the endosperm, changing it into easily digested substances, and passing it on, little by little, to the embryo as it is required.

Man also, however, has discovered the nutritive properties of the endosperm, and unceremoniously putting aside the prior claims of the embryo, he appropriates this store of food to his own use, converting it into *flour*.

**The food-stuffs present in flour.**—The value of the endosperm—or, as we may now call it, the flour—of the wheat lies in the fact that it consists almost wholly of two very important food-stuffs—starch and gluten. A little sugar is also

present. The starch and gluten may easily be separated from flour by mixing some of the latter with a little water, so as to make a stiff dough. This is tied up in an old handkerchief, and kneaded gently between the fingers in a basin of water. The starch comes through the meshes of the cloth, and turns the water milky. The kneading is continued until no further milkiness is produced, and the handkerchief is opened and the soft, india-rubbery mass of wet gluten which is left taken out. The starch settles to the bottom of the basin, and may be seen as a white deposit when the water is poured off.

**The use of yeast in bread-making.**—Yeast is added to the dough in bread-making, because when the tiny plants of which the yeast consists (Fig. 153) are supplied with a solution of sugar, and kept warm, they multiply rapidly, forming alcohol.

and giving off carbon dioxide gas. The elastic gluten which is present is blown out by this gas into little pockets, rendering the dough light and porous. During the baking process the gas is expanded by the heat, but is retained by the tough gluten. The digestibility of bread depends largely upon its light and spongy texture, so that the mere elasticity of gluten is of great use, quite apart from its value as a food.

Men have been acquainted with the properties of yeast from very early times. The ancient Greeks are known to have cultivated it, and its cells have been detected by the microscope in some bread more than four thousand years old, found in an Egyptian tomb. The Egyptians were in the habit of leaving over



FIG. 157.—A loaf of bread cut open to show the spongy nature of the interior. The large holes show that the yeast was not uniformly mixed with the dough.

a little dough or "leaven" from each baking, and they used this in subsequent bakings—the yeast which remained in the bit of dough being sufficient to leaven the new batch. Bread-making, as carried out by a woman of ancient Egypt, is described in an interesting manner in Maspero's "Dawn of Civilisation":—<sup>1</sup>

"She spread some handfuls of grain upon an oblong slab of stone, slightly hollowed on its upper surface, and proceeded to crush them with a smaller stone like a painter's muller, which she moistened from time to time. . . . The flour, made to undergo several grindings in this rustic mortar, was coarse, uneven, mixed with bran or whole grains which had escaped the pestle, and contaminated with dust and abraded particles of the stone. She kneaded it with a little water, blended with it, as a sort of yeast, a piece of stale dough of the day before, and made from the mass round cakes about half an inch thick, and

<sup>1</sup> London : S.P.C.K., 1897.

some four inches in diameter, which she placed upon a flat flint, covering them with hot ashes."

**Yeast substitutes.**—The action of yeast is not always to be depended upon, owing to occasional contamination. For this and other reasons, many attempts have been made to find substitutes for yeast in bread-making. The best-known examples of such substitutes are the various *baking-powders*. They usually consist of bicarbonate of soda and cream of tartar (p. 262), mixed with some harmless but inert substance such as starch. When the mixture is moistened with water, the cream of tartar and bicarbonate of soda act upon each other with the liberation of carbon dioxide gas. The dough made with baking-powder rises properly enough, but the bread is somewhat insipid in taste.

The "aërated bread," which is supplied at the "A. B. C." depôts in London, is made to rise by the use, in kneading, of water which has been saturated under pressure (as soda-water has) with carbon dioxide gas. The kneading process is carried out in special machines.

**Unleavened bread.**—Bread which has not been raised by yeast, or by any other gas-yielding substance, is said to be unleavened. Well-known examples are the Passover bread of the Jews (used to commemorate their hurried departure from Egypt, there being no time to put any leaven into their bread), ship's biscuit, and the innumerable fancy biscuits or "crackers."

**Changes caused by baking.**—During the baking of the bread, as has been said, the living cells of the yeast are killed by the heat, and the little bubbles of carbon dioxide gas which fill the dough are expanded; also the gluten is stiffened or coagulated, so that the bread retains its spongy texture after cooling. The outside becomes hard and dark in colour, partly owing to the drying and partly to the formation of a stiff gum (Expt. 219) in this part.

A most important change also takes place in the starch of the bread's interior. If uncooked wheaten flour is examined by the microscope, its starch is seen to consist of tiny round granules (Fig. 138). When boiling water is added to starch, its granules burst, and to a large extent dissolve in the water to form the well-known starch paste. Something similar takes place when

bread is baked. The heat bursts the granules of the starch, and changes it into a condition in which it is easily attacked by the digestive juices of the body.

**The value of bread as food.**—The chief ingredients which make up flour, and the changes which take place in them when flour is made into bread having been briefly considered, the reason why bread is so vitally important can now be understood. An average healthy man requires every day about *three-quarters of an ounce of nitrogen*. This element occurs abundantly in the air; but unfortunately the human body is incapable of making use of nitrogen unless it is supplied in the form of certain substances known as *proteids*, which can only be manufactured by living plants and animals. Some of the best-known proteids are *albumin* or white of egg, the *casein* of milk and cheese, and the *myosin* which is an important constituent of lean meat. Last of all, gluten, which is present to the extent of about 12 per cent. in good flour, is also a proteid, and is moreover easily digested. A man's daily allowance of nitrogen may consequently be obtained (p. 209) by the consumption of just over three pounds of bread.

Another very important element of food is *carbon*, and a man must be supplied every day with about  $9\frac{1}{2}$  ounces of this element in a digestible form if he is to keep up his strength. Carbon is present, not only in gluten (p. 209), but in starch (p. 221), sugar (pp. 229 and 231), and fats, all of which bodies occur in bread. In fact, good bread contains about 29 per cent. of digestible carbon, so that our "average man" will obtain his  $9\frac{1}{2}$  ounces with ease from 33 ounces of bread, and will have an ample supply of carbon if he consume the three pounds of bread required to give, in addition, all the nitrogen he needs.

White bread also contains about 1 per cent. of the mineral substances necessary for building up the bones and teeth.

The innermost of the bran-layers of the wheat-grain (marked *a* in Fig. 156) is rich in nitrogenous material—not gluten—and all the bran-layers are well provided with mineral matter. With the object of making use of these substances as food, numerous attempts have been made to introduce flours which shall contain the bran as well as the endosperm. The value of a food, however, depends so largely upon its digestibility, that



it is doubtful whether any real advantage will be obtained by including the bran, unless it can be ground much finer than it is at present.

**The use of other cereals as food.**—The grains of other cereals have, practically, the same structure as the wheat-grain, the bulk of the seed being composed of endosperm; but in no other case is the flour obtained from it so rich in easily digested nutritive ingredients. In cold countries *rye* is more easily cultivated than wheat, and it ranks next to wheat in importance. *Oats* and *barley* are not so often used for bread-making. They contain valuable food-stuffs, but are also less digestible. *Maize*, or Indian corn, *rice*, *millet*, and *buckwheat* are other examples of cereals which are used for food.

It is clear that in the endosperm of the grain of the wheat we have an almost ideal food. The ease with which this invaluable grass can be cultivated in all temperate climates makes it accessible to nearly every nation of the earth and accounts for its universal employment. Well does bread made from wheat merit the title which grateful humanity has bestowed upon it—"the staff of life."

#### SUMMARY.

**Flour** is the endosperm of the grains of wheat, rye, barley, oats, and other cereals. It consists chiefly of *starch* and *gluten*.

In **bread-making**, the flour is kneaded into a dough with water, a certain amount of yeast and sugar is usually added, and the dough is set in a warm place to "rise." The alcoholic fermentation of the sugar liberates carbon dioxide gas, which makes the dough spongy. The mass is then heated to a much higher temperature, which kills the yeast, expands the gas, hardens the gluten, bursts the starch-grains, and at last forms the *bread*, with a crisp outer crust and a firm spongy interior.

**Yeast substitutes** usually consist of bicarbonate of soda and cream of tartar. Such a mixture, called *baking-powder*, yields carbon dioxide gas on being moistened with water. *Aerated bread* is made to rise by the use, in kneading, of water saturated under pressure with carbon dioxide gas.

**Unleavened bread** is made without the use of either yeast or yeast substitutes.

**Bread as food.**—An average man's daily food must contain enough proteid matter to yield  $\frac{3}{4}$  oz. of *nitrogen*, and sufficient proteids, carbohydrates, or fats to yield  $9\frac{1}{2}$  oz. of *carbon*.

$\frac{3}{4}$  oz. of nitrogen (in the form of proteid) is contained in just over 3 lbs. of bread.

$9\frac{1}{2}$  oz. of digestible carbon is contained in 33 oz. of bread.

Bread contains about 1 per cent. of the *mineral matter* necessary for building up the bones and teeth.

### EXERCISES ON CHAPTER XIX.

1. Describe what you have observed when wheat flour is heated in an oven, and then more strongly heated over a gas flame, and finally burnt on platinum foil. What is the appearance and nature of the ash which remains? (1898.)

2. Describe the general structure of a grain of wheat. Which part is converted into flour?

3. Describe the process of bread-making.

4. What chemical and other changes do the constituents of dough undergo when it is baked?

5. What do you mean by the term *yeast substitute*? Give examples.

6. What is unleavened bread? Do you consider it as digestible as an ordinary loaf? If not, why not?

7. To what constituents does bread owe its value as a food? Briefly describe them.

8. What daily amount of bread do you consider would be sufficient for an average man, if bread were the only food he could get? Give reasons for your answer.

9. How would you determine the amount of mineral matter present in bread? What is the use of this mineral matter as a food?

10. What is leaven, and how does it act?

## CHAPTER XX

### ACETIC ACID

#### VINEGAR AND ITS PROPERTIES

**251. Appearance and acidity of vinegar.**—(a) Obtain a pint of vinegar and examine its properties.

Vinegar is generally of a brown *colour*, but “white vinegar” is colourless, like water. Notice the characteristic *smell*.

(b) By means of a glass rod place a drop of vinegar on (1) blue litmus paper; (2) red litmus paper.

The blue litmus paper is reddened; the colour of the red paper is not changed. Vinegar has *acid* properties.

**252. Effect of boiling upon the acidity of vinegar.**—Place about half a teacupful of vinegar in a porcelain dish, and heat it to gentle boiling by means of the Bunsen flame. Notice the smell of the vapour, and from time to time take out a drop of the vinegar and test it with blue litmus paper to see if it is still acid. When the liquid has nearly all boiled away, turn the flame low, and at the moment when the dish is dry take away the burner. Notice that a very small residue is left at the bottom of the dish. When the dish is cool, add about half a teaspoonful of distilled water, and stir with a clean glass rod. Place a drop of this liquid on blue litmus paper to see if it is acid.

Does the acidity of the vinegar reside in the dissolved solid matter, or is it due to something which boils away with the water?

**253. Neutralisation of acidity of vinegar.**—Wash out the dish and place in it about the same quantity of vinegar as before. To this add sodium carbonate in small pinches at a time, stirring well at each addition, until the mixture just fails to turn blue litmus red. Notice the effervescence which takes place. Try not to overshoot the mark by adding too much sodium carbonate. With care it is possible to obtain a solution which neither turns blue litmus red nor red litmus blue; that is, which is neither acid nor alkaline, but *neutral*. Carefully evaporate the neutral solution to dryness, testing at intervals, as before, to see if the liquid remains neutral during the evaporation.

Notice that a much larger residue is left in the dish than in Expt. 252. The residue in this case consists almost wholly of a salt, known to chemists as *sodium acetate*. Put the dish aside for Expt. 256.

**254. Action between acids and carbonates.**—(a) In Expt. 253 the addition of sodium carbonate to vinegar was seen to cause effervescence, due to the escape of gas. Put as much sodium carbonate as will lie on a threepenny piece into a test-tube, and add vinegar. Test the escaping gas with lime-water in the manner of Expt. 169.

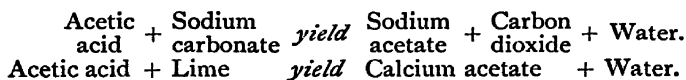
The lime-water becomes milky, showing that the gas is carbon dioxide.

(b) Repeat the experiment, using dilute sulphuric acid instead of vinegar. Carbon dioxide is again liberated from the sodium carbonate.

Vinegar has plainly the properties of an acid. It is really dilute *acetic acid*.

**255. Action between vinegar and lime-water.**—In another dish neutralise another quantity of vinegar (about as much as was used in Expts. 252 and 253) with lime-water, noticing by means of litmus paper, as before, when the liquid becomes neutral, and taking care not to overshoot the mark. Observe that in this case there is no escape of carbon dioxide gas, because lime is not a carbonate. Evaporate the neutral solution to dryness. The residue which is left is a salt known as acetate of lime or *calcium acetate*.

The production of the two salts obtained from vinegar, by the action of sodium carbonate and lime-water respectively, may be conveniently represented as follows :—



**256. Action between sulphuric acid and acetates.**—Take the dishes containing the sodium and calcium acetates, and, after noticing that each is free from smell, add a few drops of sulphuric acid to each and warm gently. Smell cautiously; the sharp, pungent odour of acetic acid is at once apparent.

The change may be represented thus :—



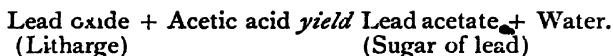
**257. Properties of acetic acid.**—Examine the properties of a sample of commercial acetic acid, and show, by a repetition of Expts. 253 and 255, that it behaves like very strong vinegar; a far larger quantity of sodium carbonate is required to neutralise half a cupful than was the case with the same volume of vinegar. Evaporate the solution to dryness, and preserve the residue of sodium acetate for Expt. 260.

**258. Action between acetic acid and iron.**—Put some iron-filings into a little acetic acid in a test-tube, and notice that a gas is given off as the iron slowly dissolves in the acid. The gas is hydrogen.

**259. Preparation of sugar of lead.**—Dissolve litharge, a compound of oxygen and lead, in acetic acid in a dish, adding litharge so long as any dissolves. Put the dish on one side, and notice that crystals soon

appear. These consist of lead acetate, or "sugar of lead," a very poisonous substance.

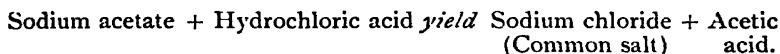
The action between the acid and the litharge may be thus represented :—



**260. Preparation of acetic acid.**—(a) Break up about half of the sodium acetate obtained in Expt. 257 into small pieces, and place them in a small retort fitted up as shown in Fig. 60. Add sufficient strong sulphuric acid to cover the salt, and allow the mixture to stand for some time. Then warm. The liquid (the "distillate") which collects in the receiver is strong acetic acid. Test a small portion by means of litmus paper, iron-filings, sodium carbonate, etc., and show that its properties correspond to those already seen to be possessed by acetic acid. Hold the receiver under the tap, if necessary, to cool the distillate to 17° C. or lower, and notice that the acetic acid then forms a solid mass of colourless crystals. When strong enough to behave in this manner, the acid is known as *glacial acetic acid*.

(b) Repeat the experiment, using hydrochloric acid instead of sulphuric acid.

In this case we have :—



**261. Acetic acid formed by alcohol and oxygen.**—Put into a watch-glass as much "platinum-black" as will lie on a threepenny piece, and form a paste by mixing it with one or two drops of pure alcohol. Put a small piece of blue litmus paper into the paste, and then cover the watch-glass with an inverted beaker, placing a split match under the edge to allow air to enter. After about twenty-four hours examine the litmus paper to see if it has changed in colour. The red colour observed shows that an acid has been formed. Warm the watch-glass gently, and notice the odour of acetic acid. The alcohol has combined with some of the oxygen of the air to form acetic acid—the platinum-black acting as an oxygen-carrier.

**262. Density of vinegar and acetic acid.**—Find the density or specific gravity of (a) vinegar, and (b) commercial acetic acid, by the method explained on p. 32.

**Acetic acid.**—Acetic acid is produced on a large scale by strongly heating wood in closed iron retorts. Water and other liquids, tar and various gases, are given off, charcoal being left in the retort when the action is finished. The gases pass on, and are sometimes used for illuminating purposes; but the watery liquids and tarry matters are collected in cooled receivers. The brown, watery layer contains, among other substances, a large quantity of acetic acid, sometimes called

*pyroligneous acid*. The acid is separated from the tar, and neutralised by lime, forming a solution of calcium acetate. This solution is distilled in order that the valuable "wood spirit," etc., present may be separated from it. It is then evaporated to dryness, and the dry residue of calcium acetate distilled with strong hydrochloric acid in copper retorts. The liquid which collects in the receiver is the crude acetic acid of commerce. It is purified by redistillation.

The strongest acetic acid solidifies between  $16^{\circ}$  C. and  $17^{\circ}$  C. to a crystalline mass. Such acid is known as *glacial acetic acid*. Pure acetic acid has a density or specific gravity of 1.055 (water = 1). It boils at  $118^{\circ}\cdot 5$  C., and mixes in all proportions with water and alcohol. Acetic acid has all the properties characteristic of acids, blistering the skin when strong, and uniting with bases to form salts and water.

**Acetates.**—An example of the action of acetic acid with one base (lime) is seen in its manufacture, when the pyroligneous acid is neutralised with lime. By its action on another base (litharge, an oxide of lead) the very important salt *lead acetate* (sugar of lead) is produced. This is extensively employed in calico-printing and in various dyeing processes. Acetic acid dissolves certain metals, hydrogen being liberated at the same time. This solvent property is utilised commercially, by dissolving iron turnings and nails in pyroligneous acid to obtain "*black-liquor*," another substance used in dyeing and printing.

When acetic acid acts upon a carbonate, carbon dioxide gas is liberated, and an acetate is formed. Thus, *sodium acetate*, which is used in medicine and as a food-preservative, is obtained by neutralising acetic acid with sodium carbonate, as in Expt. 257.

Pure acetic acid has no action on litmus, but, in the presence of water, acetic acid turns blue litmus red, as other acids do. The free acid has a characteristic sharp, pungent smell; but its salts, the acetates, are almost or quite odourless.

Pure acetic acid has the following percentage composition :—

Carbon . . . . .	40.0
Hydrogen . . . . .	6.7
Oxygen . . . . .	53.3
	<hr/>
	100.0

**The formation of acetic acid from alcohol.**—Under certain conditions alcohol may be made to unite with oxygen and become converted into acetic acid. This combination can be brought about by making finely divided platinum (platinum-black or spongy platinum) into a paste with alcohol, and leaving the paste in contact with air for some hours. The platinum acts by bringing the alcohol and oxygen into very close contact, without which the oxidation of alcohol into acetic acid could not take place; but the platinum-black does not itself undergo any change.

It has been seen (p. 234) that, when alcohol is *completely* oxidised by burning, it forms carbon dioxide and water.

**Vinegar.**—The commonest form of acetic acid is the well-known liquid *vinegar*. Genuine vinegar not only contains from 5 to 12 per cent. of acetic acid, but also very small quantities of other substances, which give it its pleasant smell and flavour—not possessed by artificial vinegar. The rest is water.

**Acetous fermentation.**—When beer, or weak wine, is ex-

posed to the air, it changes into vinegar by the combination of the alcohol with oxygen from the air to form acetic acid. The union is brought about by the activities of very minute living plants, called *bacteria*. As the growth and multiplication of these particular bacteria give rise to vinegar, they are called the *vinegar organisms*. Their cells are shown highly magnified in Fig. 158. The action of the vinegar organism in converting alcohol into acetic acid is called *acetous fermentation*, and the plant itself, like the yeast plant, is regarded as a *ferment*.

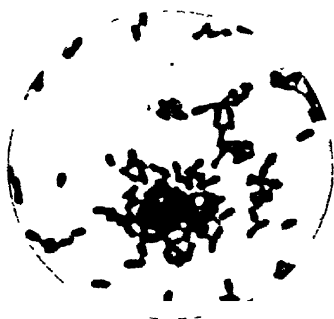


FIG. 158.—Cells of the bacterium of acetic acid (the "vinegar organism.") Highly magnified.

It should be noticed that acetic acid may be built up from starch by a series of fermentations:—

- (a) *Starch* is converted by *diastatic* fermentation into *sugar*.
- (b) *Sugar* is converted by *alcoholic* fermentation into *alcohol* and *carbon dioxide*.

(c) *Alcohol* is converted by *acetous* fermentation into *acetic acid*.

**The manufacture of vinegar.**—The cells of the vinegar organism, like those of wild yeasts (p. 235) are continually floating about as dust in the air, and so readily gain access to any exposed liquid.

In weak solutions of alcohol, containing also certain nitrogenous substances and mineral salts—which they require for food—they multiply so rapidly that one or two cells soon become many thousands, all engaged actively in converting alcohol into acetic acid.

The old method of forming vinegar consists in exposing the wine in large vats. The lids of the vats are perforated, so that the liquid is in free communication with the air.

In the *quick vinegar process*, the alcoholic liquid is kept continually trickling over beech-twigs or shavings. In this way it is broken up into small drops; and as a very large surface is thus exposed to the air, the conversion into vinegar is complete in a comparatively short time.

**Artificial vinegar.**—Diluted acetic acid coloured with caramel or burnt sugar is not vinegar, although it is often fraudulently sold as such. Genuine vinegar has been prepared by the acetous fermentation of malt-liquors or wine; and its value as a condiment is largely due to the presence of very small quantities of “extractive matters,” partly derived from these, and partly produced during fermentation.

**Uses of vinegar.**—Vinegar has been used from very early times for seasoning food, making sauces, etc. It is said to have been used by the Roman soldiers to mix with water as a beverage, and the practice is still followed by the peasantry in some parts of Southern Europe. It is now most commonly used for pickling vegetables—cucumbers, onions, gherkins, walnuts, cauliflowers, etc.

#### SUMMARY.

**Acetic acid** is a colourless liquid with a sharp characteristic odour. It is composed of carbon, hydrogen, and oxygen. It mixes with water in all proportions. Strong acetic acid, which crystallises at about 17° C., is said to be *glacial*.

Acetic acid forms salts called *acetates*, which are almost or quite odourless.



Acetic acid may be produced by the *oxidation of alcohol*. This occurs naturally, by the activities of a bacterium, the *vinegar organism*, when malt liquors or weak wines are exposed to the air. The change is known as *acetous fermentation*, and the liquid produced is *vinegar*.

*Artificial vinegar*, consisting of diluted acetic acid coloured with burnt sugar, is deficient in traces of substances which give genuine vinegar its attractive flavour.

Vinegar is used for pickling vegetables, making sauces, etc.

### EXERCISES ON CHAPTER XX.

1. What agents are concerned in the production of vinegar from weak wine or beer? By what characters is the acid of vinegar distinguished, and why does the smell of vinegar disappear when much sodium carbonate is added to the liquid? (1896.)

2. How may alcohol be obtained from grape sugar and from beer? How may it be converted into acetic acid? (1894.)

3. What acid is formed when weak wine is left exposed to the air? Explain the nature of the process which goes on. (1899.)

4. By what fermentative changes may acetic acid be derived from starch? What is a ferment?

5. Give a short account of the manufacture of acetic acid.

6. State some of the uses of vinegar.

7. Explain how acetic acid may be obtained from wood.

## CHAPTER XXI

### TARTARIC ACID

#### PROPERTIES AND USES OF TARTARIC ACID

**263. Properties of tartaric acid.**—(a) Examine a specimen of tartaric acid. It consists of colourless crystals which, when well formed, are seen to have the shape of prisms. Taste a small crystal. It is intensely sour, as most acids are. Rub a crystal, wetted with water, upon blue litmus paper. The paper turns red. Notice that tartaric acid differs from acetic acid in having no smell.

(b) *Solubility.*—Weigh a large crystal, and then powder it in a mortar. Now take a graduated jar, and, after filling it up to the 100-cubic-centimetre mark, pour water from it into the mortar in small quantities at a time. Stir well, so that as little water as possible may be used for dissolving the tartaric acid. When the acid has completely dissolved, notice from the graduations on the jar how many cubic centimetres you have used.

Remembering that a cubic centimetre of water weighs one gram,<sup>1</sup> notice that tartaric acid is soluble in less than its own weight of water.

**264. Action between tartaric acid and sodium carbonate.**—(a) Dilute the concentrated solution of tartaric acid with about three times as much water, and then add sodium carbonate in small pinches at a time until the liquid is neutral—that is, until it fails to change the colour of litmus paper. Notice the effervescence which takes place. Evaporate the neutral solution to dryness.

The residue is a salt called *sodium tartrate*.

(b) Pour tartaric acid solution upon sodium carbonate in a test-tube, and, by means of lime-water in another test-tube, test the gas which is given off, in the manner of Expt. 169. The lime-water turns milky: the gas being carbon dioxide.

The action between the tartaric acid and sodium carbonate may be represented as follows :—



<sup>1</sup> See note as to temperature on p. 31.

**265. Presence of carbon in tartaric acid.**—Place a few crystals of tartaric acid in a porcelain crucible, and heat by means of the Bunsen burner (Fig. 87). The crystals melt, and the acid becomes first yellowish, then brown, and finally black, swelling up to form a spongy mass, and giving off an odour much resembling that of burnt sugar.

The black spongy mass is largely composed of charcoal or carbon. Continue the heating, and observe that at last the mass burns completely away.

**266. Cream of tartar.**—Examine a specimen of cream of tartar. Weigh out about a teaspoonful of the crystals, and find how much water is just sufficient to dissolve them, as in Expt. 263.

A much larger quantity of water is required than is necessary to dissolve the same weight of tartaric acid.

Test the solution of cream of tartar with litmus paper. It has acid properties, turning blue litmus red.

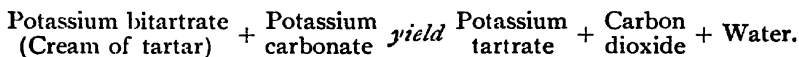
Dilute the solution to about twice its volume, and divide this diluted solution into two parts. To one part slowly add rectified spirits of wine (not methylated spirit), and stir. Before long the solution becomes cloudy, owing to the separation of cream of tartar in the solid state.

Cream of tartar is therefore less soluble in a mixture of alcohol and water than in water alone.

**267. Other tartrates.**—To the other part of the solution add potassium carbonate in small quantities at a time until the liquid is just neutral. Evaporate the neutral solution to dryness on a water-bath.

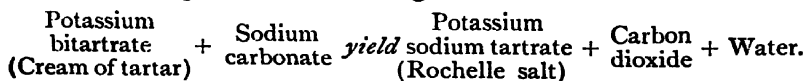
The residue is *potassium tartrate*.

Cream of tartar may be regarded as tartaric acid which has been only *half neutralised* by the addition of potassium carbonate. It is therefore known chemically as *potassium bitartrate*. When the neutralisation is completed by the addition of a further quantity of potassium carbonate, potassium tartrate is obtained. The action which takes place may be represented as follows:—



**268. Preparation of Rochelle salt.**—Dissolve as much cream of tartar as possible in about half a cupful of boiling water in a porcelain dish. Keep the solution just boiling, by means of a small Bunsen flame, and gradually add *sodium carbonate*, stirring with a glass rod, until the liquid no longer turns blue litmus paper red. Evaporate the solution to dryness.

The residue is potassium sodium tartrate—commonly known as *Rochelle salt*. The neutralisation of the potassium bitartrate has in this case been completed by sodium carbonate. Hence we may represent what has taken place in the following manner:—



**269. Argol.**—Examine a specimen of *Argol*. This is impure cream of tartar. It is generally of a reddish colour. Make a hot strong

solution, and lay it aside to cool. Crystals separate out, and most of the colouring matter remains behind in the "mother liquor." Drain the crystals; and if they are still coloured, again dissolve them in as small a quantity of hot water as possible. Allow the solution to crystallise. This time the crystals are almost or quite colourless, and consist of cream of tartar or almost pure potassium bitartrate.

Cream of tartar is prepared in this way on the large scale—by the recrystallisation of argol.

**270. Seidlitz powders.**—Make a Seidlitz powder according to the following prescription from the *British Pharmacopœia*:—

(a) Weigh out 7·77 grams of Rochelle salt (sodium potassium tartrate), and 2·59 grams of sodium bicarbonate. Both salts should be dry and powdered. Mix the two, and wrap in blue paper. This mixture is known as the *alkaline powder*, because sodium bicarbonate is alkaline.

(b) Weigh out 2·46 grams of tartaric acid, also in dry powder, and wrap in white paper. This is the *acid powder*.

The alkaline powder (in blue paper) is dissolved in nearly half a pint of cold or warm water, and the acid powder (in white paper) is then added. The draught is drunk whilst the mixture is still effervescing.

**Tartaric acid.**—Tartaric acid occurs in many fruits, especially when they are ripe. It is, for example, the chief acid in ripe grapes. The berries of the mountain ash contain considerable quantities of it, and it is found also in pine-apples, tamarinds, mulberries, cucumbers, sorrel, and other plants. All the tartaric acid of commerce is, however, derived from grape-juice.

Tartaric acid forms colourless prismatic crystals, which have a very sour taste. Like other acids, it turns blue litmus red, and neutralises bases and carbonates, giving rise to a series of salts known as *tartrates*. When only half neutralised, it forms acid salts called *bitartrates*, the best known of which is potassium bitartrate or *cream of tartar*. Tartaric acid melts at about 170° C., and on being further heated swells up into a black, spongy mass, with the separation of carbon, and giving a distinct odour of "burnt sugar" at the same time. On prolonged heating it burns completely away.

**Salts of tartaric acid.**—Tartrates and bitartrates behave somewhat similarly when heated—charring, and giving an odour of burnt sugar—but they do not burn away completely; they leave instead a residue of mineral matter. The commonest salt of tartaric acid is potassium bitartrate or *cream of tartar*. This separates, as a crust, in an impure form known as *Argol* when grape wines are stored in casks; for the salt becomes less

soluble as the amount of alcohol in the wine increases by "secondary fermentation." The argol derived from the wine of red grapes is dark-red in colour; from white grapes it is white. Purer forms of potassium bitartrate are obtained from argol by recrystallisation from hot water. When only partially purified it is known as "tartar"; but when practically all the impurities have been removed by repeated crystallisation, it is called "cream of tartar" or "cream." All the tartaric acid which comes into the market is derived from the argol which is deposited from wine.

*Cream of tartar*, or potassium bitartrate, is not a neutral salt, but has the power, like acids, of reddening blue litmus, combining with bases, and giving rise to effervescence with carbonates. It is, therefore, often called an acid tartrate. With potassium carbonate it produces neutral potassium tartrate. When it is neutralised by *sodium* carbonate it forms sodium potassium tartrate, or *Rochelle salt*, which is obtained in large transparent crystals.

**Uses of tartaric acid and tartrates.**—Both potassium bitartrate (cream of tartar) and sodium potassium tartrate (Rochelle salt) have valuable medicinal properties. They are used in the preparation of effervescing drinks. Rochelle salt is the active constituent of nearly all the so-called Seidlitz powders: the sodium bicarbonate and tartaric acid being added to produce effervescence and make the draught more palatable. Baking-powder (p. 248) generally consists of a mixture of cream of tartar and sodium bicarbonate.

Another compound of tartaric acid is the well-known *tartar emetic*. This contains antimony, and is very poisonous. It is used in medicine as an emetic, and is also employed in calico-printing.

**Composition of tartaric acid.**—Tartaric acid, like acetic acid, consists of the elements carbon, hydrogen, and oxygen, but they are combined in different proportions. Its percentage composition is

Carbon . . . . .	32
Hydrogen . . . . .	4
Oxygen . . . . .	64
	<hr/>
	100

The presence of carbon in tartaric acid is readily shown by the separation of charcoal when the acid is heated. When tartaric acid is burnt, its carbon forms carbon dioxide, and its hydrogen water, by union with oxygen.

### SUMMARY.

**Tartaric acid** is found in many ripe fruits, especially grapes.

It is composed of carbon, hydrogen, and oxygen. When pure, it forms colourless crystals which are soluble in less than their own weight of water. It has no smell.

The salts of tartaric acid are called *tartrates*. Some, such as *Rochelle salt* (sodium potassium tartrate) are neutral. Others, which are called bitartrates, have acid properties and decompose carbonates with effervescence. *Cream of tartar* (potassium bitartrate) is the commonest of these. It is obtained by recrystallisation from *argol* (impure potassium bitartrate). As the compound is less soluble in dilute alcohol than in water, argol separates as a crust in stored wines, as the proportion of alcohol in them increases.

Tartaric acid, cream of tartar and Rochelle salt have valuable medicinal qualities; advantage is taken of the effervescence caused by the first two with sodium bicarbonate in the preparation of effervescing drinks, Seidlitz powders, baking-powders, etc.

*Tartar emetic* contains antimony, and is poisonous.

### EXERCISES ON CHAPTER XXI.

1. Give an account of the properties of tartaric acid. In what respects does it differ from acetic acid?
2. What is cream of tartar? How is it prepared? What is its action upon sodium bicarbonate?
3. How do the properties of tartrates differ from those of bitartrates? Mention an example of each class.
4. Describe what is seen when (a) tartaric acid, (b) cream of tartar, is strongly heated in a crucible.
5. How would you make a Seidlitz powder? What takes place when the powder is dissolved in water?

## CHAPTER XXII

### FATS AND OILS. SOAP AND GLYCERINE

#### CLEANSING PROCESSES AND THE USE OF SOAP

**271. How to make caustic soda.**—Dissolve a quarter of a pound of washing-soda in water. Add an ounce of slaked lime. Stir well, boil for some time, and then allow to settle. The liquid consists of a solution of *caustic soda*; the sediment is chalk, or carbonate of lime. Pour off the caustic soda solution into a bottle,  *cork* it, and label it.

**272. To remove grease stains.**—Smear a piece of cloth with a bit of suet. Remove the grease by sponging the cloth with diluted caustic soda solution. Try also the effect of washing-soda on the cloth.

**273. To remove ink-stains.**—Make a *dilute* solution of bleaching-powder, and dip into it a part of an old handkerchief which has been stained with ink. Wash at once with clean water. If the stain is not removed, dip the cloth again into the dilute solution of bleaching-powder, rinsing with clean water immediately afterwards.

**274. Removal of fruit-stains.**—Stain an old handkerchief with fruit juice or jam, and dry it. Dip the stained part into dilute bleaching-powder solution, rinsing immediately with water, until the stain is removed. Fresh fruit-stains may often be completely removed by hot water alone.

**275. Effects of strong caustic soda upon cloth.**—Put a bit of calico or linen into strong caustic soda solution. Notice that the fibres of the cloth are rotted.

**276. Effects of strong bleaching-powder solution upon cloth.**—Treat a bit of cloth with strong solution of bleaching-powder. In this case also the fibres become rotten.

If caustic soda or bleaching-powder be used for removing stains from cloth, it should be in very dilute solution.

**277. The properties of suet.**—Examine a piece of ordinary beef or mutton suet. When quite fresh, it is white or slightly creamy in colour, and has a pleasant, characteristic smell. It is slightly brittle, but is difficult to break into small pieces owing to the thin membrane or skin which occurs in it. This is called *connective tissue*.

**278. Suet lighter than water.**—Throw a small piece of suet into water. It floats.

**279. The action of heat on suet.** Put a *small* piece of suet into a porcelain crucible, and heat by means of the Bunsen flame. It melts, and begins to crackle and smoke. Notice the peculiar smell. At last the fat catches fire, and burns with a very luminous flame. It burns completely away on prolonged heating.

**280. How to make soap from suet.**—Hang a small beakerful of freshly-shredded suet in boiling water until the fat is melted. To 25 c.c. of the melted fat add 75 c.c. of caustic soda solution containing 40 grams of caustic soda to the litre. Heat the mixture in a beaker surrounded by boiling water. Stir constantly until the mixture becomes pasty. Dissolve in about 100 c.c. of warm water and add 100 c.c. of saturated solution of common salt. Skim off the granular soap which separates; add cold water and stir vigorously. Skim again; wash in more cold water; skim, drain, and heat the wet granular soap in a dish over a Bunsen burner flame to drive off as much water as possible without burning the soap.

The proportions (one volume of melted fat to three volumes of caustic soda solution of the strength mentioned) are roughly those required theoretically. The quality of the soap obtained is improved by a more prolonged heating of the fat and caustic soda solution together. To increase the amount of soap obtained, the granular soap should be strained through muslin, and, during the washing process, salt solution should be added.

Such a soap, like ordinary hard soap, is chiefly composed of *sodium stearate*.

**281. Stearic acid.**—To a solution of ordinary soap in water add dilute hydrochloric acid. The liquid curdles, and white masses of stearic acid rise to the top. If the vessel is warmed, the stearic acid runs together to form an oily layer, which upon cooling again may be easily separated, rinsed with water, and examined.

**282. Oleic acid.**—Examine a sample of oleic acid, often commercially though incorrectly called “olein.” It is an oily liquid, somewhat reddish in colour, with a peculiar smell. Pour a few drops into a basin of water, and notice that they do not dissolve. Test a little of the oil with litmus paper. There is no change of colour. Stir the oil vigorously with water, and test the water with litmus paper. There is no change of colour.

Oleic acid differs from such acids as sulphuric, hydrochloric, nitric, and acetic acids in not turning blue litmus red.

**283. Make a soap from oleic acid.**—Put about a teaspoonful of oleic acid into a dry porcelain dish, and warm gently. Then add sodium carbonate in small pinches at a time, and stir with a glass rod. Notice that effervescence takes place, as would be the case with other acids and sodium carbonate (p. 253).

As more and more sodium carbonate is added, the mixture becomes thicker, until at last it is quite pasty, and on cooling becomes stiff.

This stiff mass is a *soap*, known as sodium oleate.

Rub a little of the sodium oleate soap with water on the hands. It forms a lather in the usual way.

**284. The separation of oleic acid from sodium oleate.**—Make a



solution in water of the soap prepared from oleic acid and sodium carbonate, and add dilute hydrochloric acid. Free oleic acid separates and rises to the top as an oily layer.

**285. Soap and hard water.**—Dissolve a piece of good hard soap, the size of a walnut, in a pint of warm distilled water. The solution is almost clear. Dissolve the same amount of soap in a similar quantity of "hard" water (London water answers very well), and notice that a scum collects on the surface of the solution. This scum consists chiefly of an insoluble soap of lime, *calcium stearate*.

**286. Properties of glycerine.**—It was stated (Expt. 280) that, when suet and soda are boiled together, glycerine is set free. Examine the properties of glycerine. It is a thick, colourless liquid, which flows only slowly when the vessel containing it is tilted.

(a) *Solubility in water and neutrality.*—Pour a few drops into a basin of water, and stir. The glycerine dissolves very freely in the water.

Dip red and blue litmus papers into the solution. There is no change of colour. Glycerine is neutral.

(b) *Smell and taste.*—Glycerine has no smell. Taste a drop. It is intensely sweet.

(c) *Solvent powers.*—Stir up glycerine in a dish with a small quantity of powdered litharge. The litharge dissolves. Stir up litharge with water, and notice that it does not dissolve.

Glycerine is one of the most powerful solvents known; it dissolves various substances which are insoluble in water.

(d) *Effects of heat.*—Put a few drops of glycerine into a porcelain crucible, and heat by the Bunsen flame. The liquid boils, and gives off irritating white fumes with a peculiar odour somewhat resembling that of burning fat. The fumes soon catch fire, and burn with a bluish flame as the glycerine boils away.

**Fats and oils.**—Fats and oils occur widely distributed through both the animal and vegetable kingdoms. Many of these, especially those derived from plants, are liquid at the ordinary temperature, and are known as *oils*. Others, chiefly of animal origin, which are solid at the ordinary temperature, are usually distinguished as *fats*. This is a purely artificial classification, however, as it depends merely upon differences in the melting temperatures, fats being readily liquefied by heat, and oils being solidified upon cooling.

These bodies are either colourless or pale yellow when pure. Their densities or specific gravities range from 0.88 to 0.97, and they are all practically insoluble in water. They are soluble in ether, benzene, etc. (hence the common use of the latter liquid for removing grease from gloves, etc.), and are more or less soluble in alcohol.

**Composition of beef or mutton fat.**—When suet or tallow is heated with water under a pressure from 8 to 12 times as great as that commonly exerted by the atmosphere (p. 97), or is distilled with superheated steam, it is split up into its constituents. These are (1) an organic base, called *glycerine*, and (2) the fatty acids, *stearic acid*, *palmitic acid*, and *oleic acids*. It has been seen that the result of the combination of an acid and a base is a *salt*. Beef or mutton fat is therefore to be regarded as a mixture of the three following *organic salts* :—

- (a) Stearate of glycerine, called stearin.
- (b) Palmitate of glycerine, called palmitin.
- (c) Oleate of glycerine, called olein.

The quantity of palmitin and olein in beef or mutton fat is only small. In *lard* (the fat of pigs) a much larger proportion of olein (a liquid oil) is present, which accounts for the greater softness of lard.

**Soap-making.**—Fats and oils are readily decomposed by alkalis, such as soda, potash, and lime. When, for example, suet or tallow, which consists chiefly of stearin (stearate of glycerine), is boiled with caustic soda, the soda takes the place of the glycerine, and forms stearate of soda or hard soap—the glycerine being set free. The change may be represented thus :—

STEARATE OF GLYCERINE (Stearin)	and SODA <i>yield</i>	STEARATE OF SODA (Soap)	and GLYCERINE.
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When the boiling is finished, a semi-solid mass of partly-formed soap, glycerine, and water remains. Salt is then added, and the soap rises to the top as a curd. The liquid below consists of glycerine, salt-solution, and impurities. This liquid is drawn off, and the crude soap is boiled with more soda and water until the action is complete. After being further treated to separate impurities, the soap is run off into cooling frames.

*Soft soaps* are made by boiling various oils with caustic *potash*. They contain a large proportion of water, together with all the glycerine belonging to the original oils.

**Soap as a cleansing agent.**—Soap has been used as a cleansing agent since the second century. When agitated with water it forms a “lather,” and loosens the particles of dirt by surrounding them with thin films of soap.

**Cleansing agents in general.**—The general principles underlying cleansing operations or the removal of dirt may be conveniently considered here. *Dry dust* may usually be detached mechanically, by beating, shaking, brushing, etc. *Dirt which is soluble in water* may be more or less completely removed by washing in water. Hot water acts better than cold for this purpose, as it has usually greater solvent powers. Soap, as mentioned above, assists the process by loosening the dirt. *Grease* may be converted into a soluble soap and removed by washing with washing-soda or caustic soda. Many *colouring matters*—fruit-stains, ink, etc.—are bleached, or converted into colourless substances, by the action of bleaching-powder.

In removing grease or stains from fabrics by means of caustic soda or bleaching-powder, care should be taken to use *dilute* solutions, to avoid injuring the fabric.

**The chemical nature of fats and soaps.**—The changes involved in soap-making were not understood until the beginning of the nineteenth century, when the French chemist Chevreul showed that the fats are compounds of fatty acids with glycerine. Any compound formed by replacing the glycerine of a fat or oil by an inorganic base may be called a soap.

The unpleasant scum which is formed when we wash in hard water (p. 155) is a *lime soap*. It consists chiefly of stearate of lime.

*Diachylon*, the paste which is spread on linen to form sticking-plaster, is a *lead soap* (lead oleate). It is made by heating olive oil and water with litharge (an oxide of lead). In this process the glycerine of olein is replaced by the litharge.

**Glycerine.**—Glycerine, the base of the organic salts which we call fats and oils, is a thick, colourless, neutral liquid with a very sweet taste. It is readily soluble in water, for which it has so great an affinity that when pure it will absorb half its weight of water from damp air. It is one of the most powerful solvents known, and will dissolve various substances—oxides of lead, for example—which are insoluble in water. At low temperatures glycerine becomes solid. When heated to boiling it is partially decomposed, and gives off inflammable vapours, among which is a compound very irritating to the eyes, causing a flow of tears. The same compound is produced when fat is burnt.

In Europe alone, about 25,000 tons of glycerine are annually converted into nitroglycerine, and afterwards into dynamite and other explosives.

**The acids of oils and fats.**—The acids which were originally combined with glycerine to form oils and fats may be separated from soaps by the addition of dilute sulphuric or hydrochloric acid. If, for example, dilute sulphuric acid is added to a solution in water of ordinary hard soap (sodium stearate), a white curd of stearic acid separates and rises to the top. This may be removed and examined.

*Stearic acid* is a white fatty solid at ordinary temperatures, but melts to a clear oily liquid at about  $70^{\circ}$  C. It is insoluble in water, but dissolves readily in alcohol, ether, benzene, etc. When these solutions are evaporated, the acid is deposited again in the form of crystals. Stearic acid, combined with glycerine to form stearin, is the principal acid in the fat of sheep and oxen.

*Palmitic acid*, combined with glycerine to form palmitin, is found in palm and olive oils, in suet, and in many other fats and oils. It is solid at ordinary temperatures, and melts at about  $61^{\circ}$  C.

*Oleic acid*, in combination with glycerine as olein, is the principal ingredient of olive oil, but also occurs largely in several other oils and fats. As it is liquid at ordinary temperatures, fats such as lard and goose-fat—which contain much olein—are softer than suet, which contains only a small quantity of it.

The number of fatty acids is very large, but the two of greatest commercial importance are stearic acid and palmitic acid. These are largely used, when mixed with *paraffin*, for making the so-called “stearin” candles. This mixture, on account of its comparatively high melting point and for other reasons, has been found to answer much better than tallow for this purpose.

**The value of fats as food.**—It will be seen in Chapter XXV. that fats and oils are extremely valuable foods, especially in cold countries. It is well known that the Esquimaux consume large quantities of blubber—the fat of whales, etc.—for the purpose of keeping up the vital heat.

## SUMMARY.

**Fats and oils** are **organic salts** formed from the combination of *fatty acids* (stearic, palmitic, and oleic acids, for example), with the *organic base*, glycerine. They may be decomposed into fatty acids and glycerine by superheated steam, or by heating with water under great pressure. They are insoluble in water, but soluble in benzene, ether, etc. They are lighter than water.

Fats and oils are valuable *foods*.

A **soap** is formed by replacing the glycerine of a fat or oil by an inorganic base (caustic soda, for example). Many soaps are soluble in water.

Ordinary hard soap chiefly consists of *sodium stearate*. The glycerine of stearate of glycerine (a fat) has been replaced by soda (an inorganic base).

**Glycerine**, the *base* of fats and oils, is a thick, colourless, neutral liquid with a very sweet taste. It is a very powerful solvent.

**Stearic and palmitic acids** are white, fatty solids at ordinary temperatures. Mixed with paraffin, they are used for making "stearin" candles.

**Oleic acid** is an oily liquid at ordinary temperatures.

The *fatty acids* may be separated from solutions of the soaps in which they occur, by adding dilute sulphuric or hydrochloric acid.

Of **cleansing agents**, the most generally useful is water. *Soap* acts as a cleansing agent by loosening the particles of dirt. Grease may be converted into soap and removed by using a dilute solution of *caustic soda* or *washing-soda*.

*Bleaching-powder* solution is often useful for removing stains.

## EXERCISES ON CHAPTER XXII.

1. What is the composition of suet, and what substances may be obtained therefrom by the action of caustic soda and of superheated steam respectively? (1893.)

2. Why may tallow be spoken of as a salt? Describe the two substances obtainable from it by the action of (a) water at a high temperature, and (b) boiling solution of caustic soda (1899.)

3. How is common hard soap made? What other product is formed at the same time, and what becomes of it in the ordinary process of manufacture? (1898.)

4. Compare the properties of tallow and ordinary white soap, and point out their essential differences in composition. Describe and explain the action of hydrochloric acid on a solution of soap in water. (1895.)

5. A syrupy liquid having a sweet taste may be either glycerine or a sugar solution. What experiments would you make in order to decide between them? (1897.)

6. Name seven compounds of carbon derived from vegetable or animal sources; state the source of each, and classify them into acids, salts, and neutral compounds. (1898.)

7. How would you prepare a solution of caustic soda?

8. What is the easiest method of removing grease from cooking utensils? Explain what takes place during the process.

## CHAPTER XXIII

### FLESH FOODS AND FISH. METHODS OF COOKING

#### ALBUMIN<sup>1</sup> AND ITS PROPERTIES

**287. The properties of egg-albumin.**—(a) Break a fresh egg into a basin, and separate the “white” or albumin from the yolk.

(b) *Appearance.*—Examine the egg-albumin. It is colourless and, when quite fresh, practically without smell or taste. Rub a drop between the finger and thumb, and notice its great stickiness.

(c) *Alkalinity.*—Dip a piece of red litmus paper in the albumin. The paper turns blue. Egg-albumin is slightly alkaline.

(a) *Solubility in water.*—Put a teaspoonful of albumin into a cup of water and stir gently. The albumin dissolves.

**288. Effects of heat upon albumin.**—(a) Put about a teaspoonful of egg-albumin into a clean, dry test-tube. Put a thermometer into the tube, place the tube in a beaker of water, and heat the beaker over the Bunsen flame.

The albumin remains clear until it reaches a temperature of about 58° C. At this point a slight opalescence may be seen, and at about 60° C. the albumin becomes solid, white, and opaque. This change is known as *coagulation*. Egg-albumin *coagulates* at a temperature of about 60° C.

Shake out the coagulated mass of albumin into the beaker, and continue the heating of the water. The albumin does not dissolve. Coagulated albumin is insoluble in water.

(b) Put a few drops of fresh albumin into a porcelain crucible (Fig. 87), and heat gradually by means of the Bunsen flame, noticing the changes which take place in the albumin. It first coagulates, then blackens and smokes—giving off the odour of burnt hair or feathers also noticed (p. 208) when gluten is burnt. If the heating is continued, the albumin burns almost completely away.

**289. Effects of heat on a solution of albumin.**—Stir up a teaspoonful of fresh egg-albumin in about half a pint of water. If the solution is not quite clear, filter it through filter paper in a funnel, avoiding the formation of froth on the surface of the solution. Place the solution

<sup>1</sup> Some chemists spell this word *albumen* instead of *albumin*, but the latter spelling is preferable.

in a beaker, put in a thermometer, and heat over the Bunsen flame, supporting the beaker on wire-gauze.

At a temperature of about 60° C. the solution becomes cloudy. Small particles of coagulated albumin separate as the heating is continued, and rise to the surface as a scum.

**290. No change in weight when an egg is boiled.**—Weigh an egg, and then place it in boiling water for three minutes. Take out the egg, wipe it dry and weigh again. There is practically no change in weight.

The coagulation of albumin is not due to the addition or loss of matter.

**291. The putrefaction of albumin.**—Put aside what is left of the white of egg for a few days, and notice that it soon begins to decompose, and gives off a most offensive odour. The smell of sulphuretted hydrogen, or the smell of “rotten eggs,” may easily be recognised, showing the presence of sulphur in the albumin.

**292. The properties of raw lean meat.**—Examine a small piece of raw lean beef or mutton (Fig. 159). The meat is of a red colour, owing partly to the blood which still remains in it, but partly owing to the natural colour of the fibres. It consists of a number of parallel bundles (*f*, Fig. 159) of muscle-fibre. Try to tear the bundles (*f*) apart with the fingers. It is very difficult to separate them, owing to the *connective tissue* which binds them together. The connective tissue may be seen between the bundles—especially at the cut ends—as a white or pearly glistening membrane, which is very tough.

Press a little blue litmus paper against the meat. The paper is reddened. Why?

**293. The effects of boiling upon meat.**—(*a*) Put the piece of meat into a beaker of *cold* water, put in a thermometer, and heat the beaker over the Bunsen burner, noticing the changes which take place as the water is heated. At a temperature of about 58° C. the water becomes milky, owing to the coagulation of the albumin which has dissolved out of the meat. At about 71° C. most of the coagulated albumin rises to the surface as a brownish scum, leaving the liquid much clearer. Let the water simmer gently for half an hour, and then take out the meat and examine it.

The colour of the meat has changed. It is no longer a bright red, but is brownish and much paler in colour.

Try now to tear apart the muscle-bundles. They separate much more easily than before boiling.

(*b*) Take a similar piece of meat and plunge it at once into *boiling* water. Notice that the water does not become cloudy as in the last experiment, and that the scum which rises to the top is much smaller in amount. Let the water simmer for a few minutes, and then lower the Bunsen flame and keep the water at a temperature of 80° or 85° C. for half an hour. Take out the meat and examine it.

The outside is firm, and forms a sort of crust. Cut the meat open, and notice that the inside is juicy.

**294. Amount of water in lean meat.**—Weigh a dish, and cover the bottom evenly with small pieces of lean beef. Weigh again. The



difference in weight is the weight of beef taken. Dry in the air-oven for about three hours at a temperature of from  $100^{\circ}\text{C.}$  to  $105^{\circ}\text{C.}$  Cool, and weigh again, and calculate the percentage of water lost during drying.

Example :—

Weight of dish + undried beef	... ..	=	$39\cdot38$	gms.
„ empty dish	... ..	=	$32\cdot34$	„
„ beef taken	... ..	=	$7\cdot04$	„
Weight of dish + undried beef	... ..	=	$39\cdot38$	„
„ „ + dried beef...	... ..	=	$34\cdot24$	„
• „ water lost during drying	... ..	=	$5\cdot14$	„

$$\therefore \text{Percentage of water in the beef} = \frac{5\cdot14 \times 100}{7\cdot04} = 73\cdot01.$$

**295. Amount of mineral matter in lean meat.**—Strongly heat the dish over the Bunsen burner. Notice the smell of the gases evolved. Continue the heating until only a white ash is left in the dish. Cool and weigh, and calculate the percentage of mineral matter in the beef originally taken.

**296. Amount of water and mineral matter in fat meat.**—Repeat the last two experiments, using fat meat. Notice the behaviour of the meat during heating, and calculate the percentage of water and mineral matter present.

**Albumin or white of egg.**—The egg of a bird consists chiefly of shell, “white” or albumin, and yolk. The albumin when fresh is a sticky, colourless liquid, practically without taste or smell, slightly alkaline, and soluble in water. When heated, it becomes opaque at a temperature of about  $60^{\circ}\text{C.}$ , and forms a white solid which is insoluble in water. On continued heating it blackens, and gives off smoke smelling strongly of burning hair or feathers. We have seen that a similar smell is produced when gluten (p. 208) and lean meat are burnt. In each case the peculiar odour is due to compounds of *nitrogen* which are liberated.

Albumin, gluten, and the *myosin* of lean meat are very similar in chemical composition (p. 302), and each belongs to the class of bodies which chemists call *proteids*—foods which are absolutely necessary for the maintenance of animal life, inasmuch as animals are unable to make use of nitrogen from any other source than proteids.

In general the proteids of animal origin, such as those contained in meat, fish, eggs, etc., are more readily digested than

proteids formed by plants. There are various reasons for this. One is that animal foods are already in a condition more nearly approaching the composition of our own bodies, and therefore require less change before they are ready for use. Another reason is that the food materials of vegetables are often enclosed in tough-walled cells, which the digestive juices cannot easily break down (Fig. 162). As a consequence, much of the nutritious matter of vegetables is often either wasted, or it throws an undue strain upon the digestive organs unless great precautions are taken in cooking.

### The structure of meats.

—The "meat" of animals consists of the lean or muscular tissue, and the fat which occurs in association with it. The latter may be present in quantities too small to be detected, except by special methods, as is the case

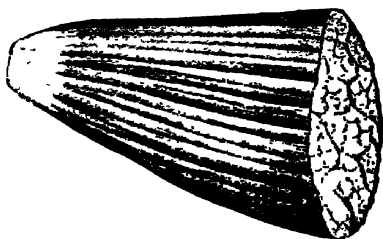


FIG. 159.—A piece of muscle (lean meat) showing the bundles of muscular tissue.

with veal, young chickens, etc., or it may form masses of appreciable size. The fat should be firm and white. The lean of meat consists of bundles (Fig. 159) of muscular tissue bound together by tough thin sheets of *connective tissue*. The bundles in their turn are composed of very fine *muscular fibres*, too small to be seen by the naked eye. These are also bound together by connective tissue.

**The composition of lean meat.**—From 50 to 75 per cent. of lean meat usually consists of water (Expt. 294). The rest consists partly of proteids, partly of fat, and partly of mineral matter. When the meat is burnt, this mineral matter remains as ash.

The average proportion of proteid matter present is about 20 per cent. The chief proteid present in meat is *myosin*. It consists, like other proteids, of carbon, hydrogen, oxygen, and *nitrogen*, with traces of sulphur. It is present in lean meat in a semi-solid gelatinous condition.

**The composition of fish.**—Generally speaking, the com-

position of fish is very similar to that of meat, although fish usually contains a smaller proportion of fat. There is, however, great variation in this respect among different kinds of fish, some being noticeably oily,\* while others are as free from fat as veal or rabbit. Like meat, fish is chiefly valuable on account of the proteids which it contains, and fish and lean meat are about equally digestible. The old idea that fish is an especially valuable food for brain-workers is now discredited by physiologists.

**Differences between animal and vegetable foods.**—The composition of vegetable foods differs largely from that of meat and fish. The flesh foods are rich in proteids and fats, but are almost destitute of carbohydrates (p. 221). Most of the cereals (p. 245) and other vegetable foods are deficient in proteids and fats, but contain large quantities of carbohydrates.

An important exception is found in peas (Fig. 165), beans, lentils, and the seeds of other leguminous plants. These contain about as much proteid matter as meat, and differ chiefly in *composition* from meat by the substitution of carbohydrates for fats.

Flesh and vegetable food-stuffs are thus supplementary, each supplying what the other lacks. Their special and comparative values as foods will be more fully considered in Chapter XXV.

**Reasons for cooking meat.**—Although raw meat is probably more digestible than cooked, it is customary among all civilised nations to submit it to a process of *boiling, stewing, roasting, or frying* before eating it. One of the chief objects of cooking meat is to soften and loosen the tough connective tissue which binds its various parts together. The meat is thereby brought into a condition in which it is more completely acted upon by the digestive juices of the stomach, intestines, etc. At the same time new and appetising flavours are developed in the food, and its appearance is rendered more inviting by the changes in colour, etc., which are brought about during cooking. These changes indirectly assist digestion by causing a greater activity of the digestive organs. A more important object of cooking is the destruction of bacteria and other *parasites* which may be present in the meat. Unfortunately, the ordinary methods of cooking are generally quite insufficient for this purpose, as the interior of a large piece of meat is seldom heated to a temperature high enough to kill all forms of life.

**The effects of boiling upon meat.**—Whether meat is put into cold or boiling water in this method of cooking depends upon the object which the cook has in view ; for the result in the two cases is very different.

If a rich broth or soup is desired, the meat is *placed in cold water*, which is gradually raised to the boiling point. As a result, the water dissolves much of the soluble albumin, the organic salts, and the flavouring substances which are present in the meat. As the temperature rises, more and more of these bodies are extracted. Small quantities of lactic acid are also formed, and these assist in the solution of certain constituents of the meat. When the water reaches a temperature of about  $58^{\circ}$  C., the dissolved albumin begins to coagulate, and at about  $71^{\circ}$  C. it separates into small particles which rise to the top of the broth, leaving the liquid below comparatively clear. Further heating coagulates the insoluble proteids which are still present in the meat, and also changes some of the connective tissue into *gelatin*, which dissolves in the broth. By this treatment the meat is left nearly tasteless, but it is nevertheless very nutritious, and should not be rejected. It still contains the greater part of the proteid matter which was originally present.

If the object of the boiling is to obtain juicy and well-flavoured meat, the piece should be *plunged at once into boiling water*. The result of this treatment is to coagulate the albumin on the outside of the meat, and thus form a sort of crust, which protects the inside from the solvent action of the water. The water should be kept at the boiling point for about ten minutes, so that the pores of the meat may be thoroughly sealed by the coagulation of the albumin. It is not, however, advisable to maintain this temperature, as the albumin of the interior would thereby be also coagulated and the meat rendered hard. It is better to let the water cool to about  $80^{\circ}$  C., and complete the cooking at this temperature. Although a longer time is thus required, the result is a much more tender joint.

**The boiling of fish.**—A similar principle is adopted in boiling fish, but as the active boiling of the water would tend to break up the fish, the latter is put into water which does not quite boil. After the albumin has coagulated, the cooking is completed at a temperature of about  $80^{\circ}$  C.

**Stewing.**—The value of this method of cooking meat does not appear to be fully recognised. It combines the advantages of the two methods of boiling described above, as it yields not only a rich broth but also juicy and tender meat. The meat is cut into small pieces, and placed in cold water. The water is then gradually heated to about  $80^{\circ}$  C., and this temperature is maintained for three or four hours or even longer.

**Extracts of meat.**—Ordinary broth and beef-tea are by no means so nutritious as they are generally supposed to be. Their value is largely due to minute quantities of extractive matters, which are strongly flavoured, and have a stimulating effect upon the nervous system.

**Beef-tea** is made by putting lean, juicy meat, cut into small pieces, with cold water into an earthenware jar. This is put into a saucepan of water, which is allowed to simmer gently from three to five hours. The liquid is strained through a coarse sieve, and set aside to cool.

True *meat extract* contains neither albumin, gelatin, nor fat; and its strong taste is due merely to flavouring matters and mineral salts. It is not a food, but a *stimulant*, resembling tea or coffee in its action.

The yield of lean beef to water is only 2 per cent., and even very strong soups, prepared as described on p. 277, with the addition of herbs and flavouring matters, contain 96 per cent. or more of water. A large quantity must, therefore, be consumed, if soup is to supply much nutriment to the body.

**Steaming.**—In steaming, the food is surrounded by steam instead of by water. The changes which take place are essentially the same as those brought about by boiling.

**Cooking by dry heat.**—Roasting, frying, or broiling is a more economical method of cooking meat than boiling, as the juices are saved. As in boiling, the coagulation of the albumin which takes place especially on the outside, has a tendency to make the meat more or less hard and difficult of digestion; but the flavours, which are developed at the same time, perhaps counteract this disadvantage to a large extent, by stimulating a freer flow of digestive juices.

Part of the juice of the meat is driven out by the heat during roasting as in boiling. This is, however, saved and used as gravy with a roast, whereas it is lost with boiled meat.

It should be noticed that the most obvious difference between roasting, steaming, and boiling lies in the medium in which the meat is cooked. In roasting, this is hot air ; in steaming it is steam ; and in boiling it is hot water. The difference is more apparent than real, for, in each case, when the cooking is properly carried out, the meat is really cooked in its own juice.

A small joint should be roasted at a higher temperature than a large one. Meat conducts heat so slowly that if a very hot fire is used for cooking a large piece of meat, the outside is burnt to a cinder before the piece is warmed through. On the other hand, it is best to employ fairly intense heat for cooking a steak or chop. The outside albumin is quickly coagulated, stopping up the pores and keeping in the juice ; and the inside is thoroughly cooked before the outside has had time to burn.

**Method of cooking decided by composition.**—The various “joints” and “cuts” of an animal differ considerably in composition, and this must be taken into account in deciding the method of cooking to be adopted. Juicy and tender pieces are generally roasted. The harder and tougher pieces are stewed, as continued heating at a moderate temperature is best adapted for softening the tissues. Pieces containing much gristle require either stewing or slow boiling ; dry heat simply hardens the gristle, and makes it useless for food.

**Composition of various cuts of meat.**—The principal “joints” of beef

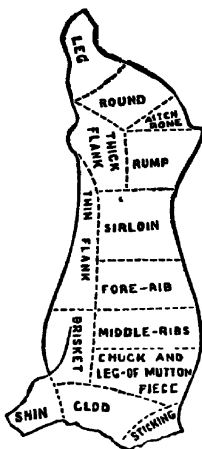


FIG. 160.—Joints of beef.

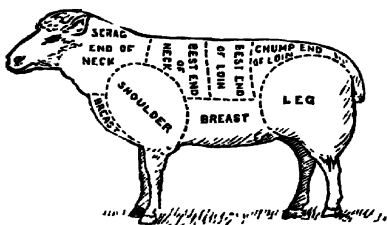


FIG. 161.—Joints of mutton.

and mutton are shown in Figs. 160 and 161. Their value depends not only on the amount of bone, gristle, and fat which are present, but also on the tenderness of the meat itself.

Of *beef*, the piece which contains the most bone is the shin. Fifty per cent. of bone is not an unusual proportion. The price of shin of beef ( $3\frac{1}{2}$ d. to 4d. a pound) is correspondingly low. *Rump steak*, on the other hand, which contains no bone, sells readily at 1s. a pound. The amount of fat also varies considerably in different regions of the body. The sirloin, ribs, flank, and brisket generally contain most fat—40 per cent. in some cases.

In *mutton*, which is sold with the natural bone belonging to each cut, a similar variation of composition may be observed. The greatest proportion of fat occurs in the loin and the lower or "best" end of the neck.

*Veal* is comparatively free from fat, while *pork* often contains a very large quantity.

**Variation due to age, etc.**—The composition of the same cut may vary considerably in two animals of the same species, according to age, sex, breed, previous feeding, and other circumstances.

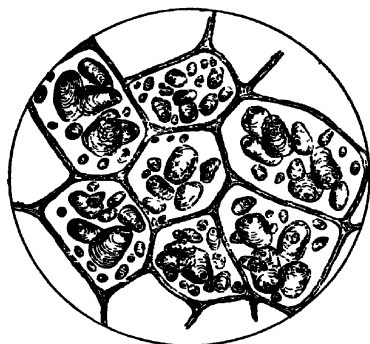


FIG. 162.—Cells of a raw potato, with the starch grains in the natural condition. (Highly magnified.)

**Reasons for cooking vegetables.**—It has been said (p. 276) that cooked meats are, strictly speaking, less easily digested than raw meats. This is largely due to the coagulation of the albumin which is brought about during cooking.

Vegetable foods are comparatively poor in proteids, so that in their case this objection does not hold good. On the other hand, the special value of vegetables lies in the carbohydrates—particularly starch—which they contain.

The food-stuffs present are enclosed in cells, which have often thick and tough walls, impervious to the digestive fluids. This is well seen in Fig. 162. On cooking, the contents of the cells swell to such an extent that the walls are burst open (Figs. 163 and 164). The insoluble starch is at the same time changed into a pasty and soluble form (pp. 217 and 248).

**The cooking of peas, beans, etc.**—Peas (Fig. 165), beans, lentils, and the seeds of leguminous plants generally,

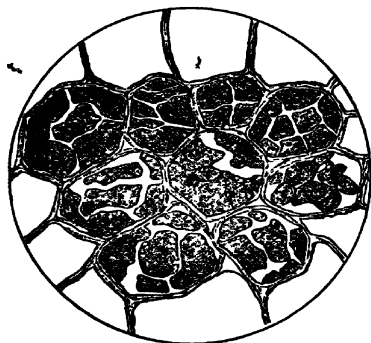


FIG. 163.—Cells of a potato boiled in water for half an hour. (Highly magnified.)

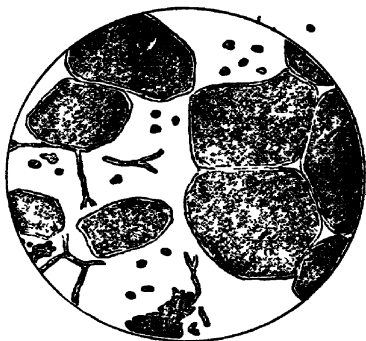


FIG. 164.—Cells of a potato well steamed and mashed. (Highly magnified.)

are remarkable for the large amount of proteids which they contain. On this account they may be used to a great extent as a substitute for animal foods. The objects of cooking these seeds are to soften and break up the cell-walls, to make the proteid ingredient (called *legumin*) digestible and attractive in flavour, and to burst the starch grains. These results are best obtained by long, slow cooking in *soft* water. The use of hard water is especially to be avoided, as lime salts form an insoluble compound with the proteid legumin.

#### SUMMARY.

Fresh **egg-albumin** is soluble in water. At 60° C. it coagulates, becoming white and opaque, and insoluble in water. It is a *proteid*.

The **myosin** of lean meat and fish is very similar in composition to egg-albumin.

Meat is **cooked** (1) to soften its connective tissue; (2) to make it more appetising; (3) to kill bacteria and other parasites.

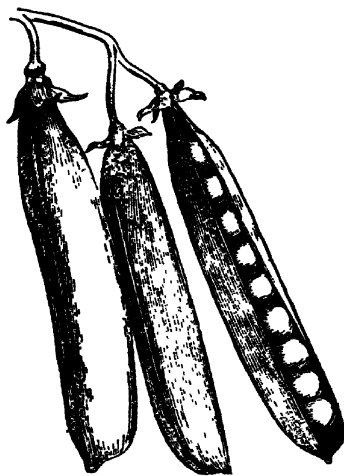


FIG. 165.—Pods and seeds of the garden pea.



**Boiling.**—When meat is put into *boiling* water, the outside albumin is coagulated, forming an impervious crust which keeps in the juices.

When meat is put into *cold* water, the temperature of which is gradually raised to the boiling point, the soluble matters are dissolved out into the broth before the albumin coagulates.

In **stewing**, the meat is put into cold water, and the cooking carried out at 80° C.

In **roasting, broiling, and frying** the juices of the meat are retained. A small joint should be roasted at a higher temperature than a large one.

**Vegetables** are cooked (1) to soften and break up the cell-walls in which the food materials are contained; (2) to burst starch grains; (3) to develop appetising flavours.

Animal foods contain practically no carbohydrates, but are rich in proteids and fats.

Most of the cereals and other vegetable foods are deficient in proteids and fats, but contain large quantities of carbohydrates.

**Peas, beans,** and other leguminous seeds contain about as much proteid matter as meat does. They should be cooked in soft water.

### EXERCISES ON CHAPTER XXIII.

1. Describe the properties of white of egg, and the changes which it undergoes on heating.
2. Compare the composition of meat, fish, and lentils.
3. What are the objects of cooking vegetables? Why are vegetable foods considered less digestible than animal foods?
4. What is the essential object of cooking processes? Explain the changes which meat and bread respectively undergo when baked. (1893.)
5. Explain the changes which meat undergoes in cooking, and indicate the essential differences between the processes of boiling and stewing. (1896.)
6. How should beef-tea be made? What food substances does it contain, and what value has it as a dietetic? (1897.)
7. How is meat changed by the process of roasting? What rules would you observe in roasting a joint? (1895.)
8. Explain the differences between *boiling, roasting, and stewing* meat. (1899.)
9. What are the characteristics of good meat? Explain how to make a good meat stew. (1897.)
10. Which is the more nutritious, rice or pea-flour? Upon what do their relative qualities depend? (1894.)

## CHAPTER XXIV

### MILK, BUTTER, AND CHEESE

#### THE PROPERTIES AND USES OF MILK

**297. New milk: appearance.**—Examine new milk. It is a white opalescent liquid, which often has a slightly yellowish tinge.

**298. Microscopic appearance.**—Place a drop of milk on a clean microscope-slide, cover it with a cover-slip, taking care not to include any air-bubbles, and examine with both low and high powers. Small rounded globules of fat are seen. Milk owes its white colour to the myriads of fat globules which are diffused through it. When fat is thus broken up into minute separate particles, it is said to form an *emulsion*.

Milk is a natural emulsion.

**299. The formation of cream.**—Allow some new milk to stand overnight in a glass vessel. The fat rises, and a layer of cream forms on the top of the milk; the line of demarcation between the cream and the milk below is very clear. The cream occupies about 7 per cent. of the whole volume. Carefully remove the cream with a spoon, and preserve the *skim milk* which is left for Expt. 301.

**300. Density of new milk.**—By the method explained on p. 32 find the density or specific gravity of new milk. It will probably lie between 1.029 and 1.034. Make a note of this.

What is the weight in lbs. of a gallon of the milk?

**301. Density of skim milk.**—When the lighter constituent of a mixture is removed, the remainder has a higher density or specific gravity than the original mixture. Determine the specific gravity of skim milk—that is, milk from which the light cream has been removed.

Skim milk is noticeably denser than new. Its specific gravity is about 1.035.

**302. Density of watered milk.**—As milk is denser than water, the addition of water to it lowers its density or specific gravity. Add about a tenth of its volume of water to a sample of new milk, the specific gravity of which has already been found. Stir, and find the specific gravity of the mixture. It is less than that of the unwatered milk.

As the specific gravity of milk is raised by the removal of its cream, and lowered by the addition of water, it is possible first to remove some of the cream and then add water in such proportions that the resulting specific gravity is the same as that of the original milk.

The specific gravity alone is therefore not a trustworthy test of the genuineness of milk.

**303. Amount of solids in milk.**—Measure with a graduated pipette 10 c.c. of a sample of milk of known specific gravity,<sup>1</sup> and run this into a small, previously weighed dish. Evaporate the milk to dryness on the water-bath, and drive off the last traces of moisture by afterwards heating the dish with its residue in the air-oven at a temperature of between 100° C. and 102° C. for two hours. After cooling, weigh the dish and residue, and calculate the percentage of solid matter contained in the milk taken.

Genuine milk usually contains about 12.5 per cent. of solid matter.

**304. Mineral matters in milk.**—Strongly heat the dish with its residue over a small, colourless Bunsen flame. The residue blackens, smokes, and bursts into flame. On prolonged heating, it all burns away, with the exception of a little white ash, consisting of the mineral or inorganic constituents of the milk.

**305. Milk curdled by a weak acid**—Add a little vinegar to half a cupful of milk. The milk curdles, separating into two portions—a thick clot or curd, and a thin watery liquid.

**306. Milk curdled by rennet.**—To half a cupful of milk warmed to blood-heat (98° F.) add about half a teaspoonful of “concentrated essence of rennet,” which may be obtained from the chemists. Stir, and then cool the mixture. The milk has curdled. Break up the curd with a glass rod, and filter off a little of the watery liquid or whey.

**307. The presence of milk-sugar in whey.**—Add about a teaspoonful of whey to twice as much Fehling’s solution (p. 214) in a test-tube, and put the test-tube into boiling water for five minutes. A red precipitate shows the presence of a sugar. In this case the sugar is lactose or milk-sugar (p. 229).

**308. The souring of milk.**—Put aside a small dishful of milk for a few days. The milk turns sour owing to the conversion of lactose into lactic acid. The lactic acid formed curdles the milk as the vinegar did in Expt. 305.

**309. How to make butter.**—Put a cupful of cream into a jug, and beat it with a wooden spoon. After a short time the fat of the cream sticks together to form lumps of butter. Take these out, and knead them in a porcelain dish until as much buttermilk as possible has been squeezed out.

**310. Properties of butter.**—Examine the properties of ordinary butter. It may vary in colour from white to yellow. It has a characteristic smell and taste. Fill a small beaker with butter, and put the beaker in a warm place—on the top of the air-oven, for example—until the butter is completely melted. It then forms two separate layers: an upper layer of butter fat with a little curd; and a lower

<sup>1</sup> The weight of milk taken is therefore known.

smaller layer consisting of water, curd, and salt. Pour off the upper layer into a dry filter, and collect the clear butter fat which runs through.

**311. The Valenta test for pure butter fat.**—Graduate a test-tube by running in 3 c.c. of water from a burette, and making a scratch with a file upon the tube at the level of the surface of the water. Run in another 3 c.c., and make another scratch at the 6 c.c. level. Pour the water out, dry the test-tube thoroughly, and pour in 3 c.c. of the strongest acetic acid. Then pour in 3 c.c. of the melted butter fat. Cork and shake the tube, and put it into water and heat the water. Notice the temperature at which the contents of the tube remain clear after shaking.

Equal volumes of strong acetic acid (containing 99 per cent. or more of pure acetic acid) and pure butter fat form a clear liquid at a temperature of about 40° C.

**312. The behaviour of margarine fat with acetic acid.**—Obtain a sample of ordinary margarine, and treat it in the same way: melting and filtering off the clear fat, and heating 3 c.c. of it with 3 c.c. of strong acetic acid. A much higher temperature (75° C. or higher) is required to give a clear liquid than is the case with pure butter fat.

**Milk.**—Milk is the fluid secreted by the females of all mammalian animals for the nutrition of their young. It is the most perfect food in Nature.

Cows' milk, which is most largely consumed, is an opalescent liquid, either white or with a slightly yellowish tinge. It has a mild, sweet taste, and a faint, pleasant odour. Its colour is due to the presence of very fine globules of fat, which may be distinctly seen when a drop of milk is examined under the microscope. These fat droplets vary in size, but are usually not more than  $\frac{1}{2500}$  of an inch in diameter. They are kept from running together by the thin film of milk which surrounds each droplet.

A liquid like milk, which contains finely divided particles of fat in this condition, is called an *emulsion*.

When new milk is allowed to stand, the particles of fat, which are lighter than the liquid in which they are suspended, rise to the top, and form a layer of *cream*. When the cream is removed, either by skimming or by one of the "separators" devised for the purpose, the liquid remaining is known as *skim* or *separated* milk.

The density or specific gravity of new milk varies between 1.029 and 1.034. That of skim milk, from which the light fat has been removed, is of course greater, being about 1.035. The addition of water to new milk, on the other hand, lowers the specific gravity. It is possible, by first removing some of the cream and afterwards adding water, to obtain a liquid having

the specific gravity of genuine new milk. The specific gravity alone, therefore, cannot be relied upon as a criterion of the genuineness of milk.

**Curds and whey.**—When milk is allowed to stand for a considerable time, it becomes *sour*, owing to the formation of lactic acid. The action of the acid upon the constituents of the milk causes it to separate into two portions: (1) a clot or *curd*, consisting chiefly of the fat and a proteid substance known as *casein*; and (2) a watery portion, the *whey*. This change may also be brought about by the direct addition to the milk of a weak acid, such as vinegar.

**Rennet.**—The separation of milk into curds and whey is effected in cheese-making (p. 290) by a ferment called *rennet*, which is extracted from the fourth stomach of a calf. Its action is particularly interesting, because the same ferment is present in the stomach of man, and brings about the same change, as a preliminary to digestion, in milk used as human food.

**The composition of milk.**—There is some variation in the proportions of the constituents occurring in genuine cows' milk, but the average composition of 120,540 samples, analysed by Dr. Vieth, was as follows:—

Total Solids	{	Fat . . . . .	4·1 per cent.
		Solids not fat . . . . .	8·8   "   "
		Water . . . . .	87·1   "   "
			<hr/> 100·0

Thousands of analyses made by other chemists, notably H. Droop Richmond, Dr. Vieth's successor in the laboratory of the Aylesbury Dairy Company, fully confirm these results.

For some years the low standard adopted by the Inland Revenue authorities, apparently founded upon the composition of the liquid yielded by certain abnormally ill-fed cows, has made it difficult to restrict properly the sale of adulterated milk. The standard of the Society of Public Analysts is as follows:—

Total Solids	{	Fat . . . . .	3·0 per cent.
		Solids not fat . . . . .	8·5 " "
		Water . . . . .	88·5 " "
			<hr/> 100·0

Even this is to be regarded as the composition of very poor, but *possibly* genuine, milk. It is hoped that a standard more nearly representing milk of average quality than that hitherto in force may shortly be made legal.<sup>1</sup>

The *fat* of milk, which is the source of butter and an important constituent of cheese, consists of a mixture of different fats. These are compounds of glycerine with certain organic acids: palmitic, oleic, and butyric being the chief.

*Casein* is the chief proteid (p. 210) contained in milk. It is thrown down as a curd when dilute acid or rennet is added. A small amount of *albumin*, somewhat similar to egg-albumin (p. 274), also occurs in milk. About one-quarter of the solid matter of milk consists of proteids.

*Milk-sugar*, or lactose, is the principal carbohydrate of milk. It has the same composition as cane sugar (p. 229), but is much less sweet. About 5 per cent. of milk consists of milk-sugar. Its presence may be readily shown in whey; that is, in milk from which the casein and fat have been removed by means of rennet. The whey is heated with Fehling's solution (p. 214), when a red precipitate is obtained.

When milk is kept in a warm place for some time, its milk-sugar is acted upon by certain minute plants or bacteria (Fig. 166), and decomposed. Lactic acid is formed, and curdles the milk.

A small quantity of *mineral matter* is also contained in milk. This is left as a white ash when milk is evaporated to dryness and the residue ignited.

**Human milk.**—Human milk contains the same constituents as cows' milk, but in somewhat different proportions. Human milk is richer in sugar, but contains a smaller quantity of proteids than cows' milk. It is therefore customary, in



FIG. 166.—Lactic acid bacteria, very highly magnified. (After Pasteur.) In order to give an idea of the size of the bacteria, some yeast-cells are figured among them.

<sup>1</sup> During the passage of this book through the press the standard of the Society of Public Analysts has been made legal, coming into force on September 1st, 1901.

preparing cows' milk for the use of infants, to dilute it with water and add sugar.

**The value of milk as a food.**—It should now be clear why milk has, from the earliest times, been considered so valuable a food. Not only are all the classes of nutrients—proteids, fats, carbohydrates, and mineral matters—contained in it, but they are also present in more nearly the required proportions than in any other single food. A quart of milk contains about as much nutriment as 12 ounces of beef or as 6 ounces of bread; but the food materials of the milk form a better-balanced diet than the bread, and those of the bread in their turn are more suitably combined than those of the beef. The chief disadvantage of milk as an entire food for adults lies in the large proportion of water present. The proportion of proteid matter, as compared with the fat and carbohydrates, is also slightly higher than is necessary. Milk is, however, especially suitable for mixing with other foods, as, for example, the cereals (Chapter XIX.). The nutritious qualities of bread are very materially improved if milk is used instead of water in mixing the dough.

**Skim or separated milk.**—Milk which has been skimmed by hand is now very rarely met with, as the cream is almost always removed by means of one of the various mechanical separators. The vessel containing the milk is made to rotate rapidly, and the light cream collects at the centre and is drawn off. The best separators leave only a trace of fat in the milk. The proteids and milk-sugar which still remain are, however, very useful foods, and give skim milk a value which does not appear to be fully recognised.

**Cream.**—Cream is obtained from milk, either by the use of the separators referred to above, or by warming the milk in pans for several hours. In the latter case the cream, which rises to the top, is very firm. It comes into the market as "Devonshire" or "Cornish" cream.

Cream should contain about 45 per cent. of fat. It is the fat which gives the characteristic "thickness" to genuine cream, but cream is sometimes sold which has been artificially thickened by the addition of gelatin. Cream turns sour so rapidly that a preservative is generally added before it is put on the market. This is commonly a mixture of borax and boracic acid.

The value of cream as a food depends upon the fat which it contains, but the price is generally so high that its use can scarcely be regarded as economical.

**Condensed milks.**—Condensed milks, which are now so common, usually consist of milk which has been concentrated to one-third its original volume. When this is the case, it is merely necessary to add twice the volume of water to obtain a fluid having the composition of the original milk.

Many brands contain added sugar, and it is not uncommon to find on the vessels containing these a recommendation to dilute with an amount of water which would yield a liquid considerably weaker than genuine milk. It is obvious that great danger is involved in using the weak fluids thus obtained as the chief or only food of young children.

The so-called “condensed skim milks” are to be avoided, especially for the purpose for which they are generally used—the feeding of infants. They contain practically no fat whatever, and are quite unsuited for the nourishment of children.

**Butter.**—When cream is shaken vigorously, the particles of fat join together to form butter, and the milk which was entangled with them is left behind as buttermilk. When properly treated, a quart of good cream yields about a pound of butter.

Butter usually contains about 85 per cent. of fat. The remainder consists of water and curd (casein), with a varying quantity of added salt. Properly made butter contains from 12 to 16 per cent. of water. When more than this proportion is present, the excess has generally been purposely added as an adulteration.

The *fat* of genuine butter has of course the same composition (p. 287) as that of milk. It may be separated from the other constituents by melting the butter. The water, salt, and most of the curd sink to the bottom of the vessel, and the fat may be filtered through a dry filter-paper in a funnel.

**Margarine**, which is made by churning beef or mutton fat (p. 267) with milk or cream, to give it the flavour and smell of genuine butter, is a cheap substitute for butter, and is often fraudulently sold as butter. It consists chiefly, like suet, of compounds of glycerine with stearic, palmitic, and oleic acids. It differs from butter principally in containing no *butyrin*, or



butyrate of glycerine, which is an essential constituent of butter fat.

A readily applied test of the genuineness of butter consists in mixing equal volumes of the clarified fat and the strongest acetic acid, and heating the mixture. With genuine butter the mixture becomes clear at about 40° C., while with the fat of ordinary margarine a temperature of 75° C. or more is required. "Nut margarine" yields a clear mixture at ordinary temperatures.

It should be pointed out that well-made margarine is a perfectly wholesome article of food. It is to be objected to only when it is sold as butter, which commands a higher price.

**Cheese.**—In the manufacture of cheese, milk is curdled, generally by means of rennet, and the curd is cut up into small pieces and allowed to stand until it has "ripened." Characteristic flavours are produced during the process of ripening. The change is due to the action of certain bacteria and moulds. When the ripening is complete, the curd is placed in a press, and the excess of water is removed by a gradually increasing pressure.

Cheese consists of proteids—chiefly casein—fat, water, and a little mineral matter. The proportions of these naturally vary with different kinds of cheese.

The only adulteration of cheese which is practised consists in the substitution of margarine fats for butter fat. Such "margarine cheeses" are not often exposed for sale in this country.

#### SUMMARY.

**Milk** is a dense, white fluid, containing about 13 per cent. of solid matter, which consists partly of fat and partly of non-fatty compounds.

The **fat** is suspended in the liquid in minute drops, forming an *emulsion*.

The **solids not fat** are in solution. They consist chiefly of *casein* (a proteid), *milk-sugar* (a carbohydrate), and a small quantity of *mineral matter*.

On standing, the fat globules rise to the surface as a layer of *cream*. When this is removed, the liquid left is known as *skim* or *separated milk*.

The best brands of *condensed milk* are prepared by evaporating new milk to one-third its volume. Added sugar is often present.

When *rennet* or a weak acid is added to milk, the fat and casein separate in the form of *curd*. The milk-sugar remains in the liquid or

*whcy.* Curd produced in this way is made into *cheese*, by first allowing it to ripen, and then pressing it to remove excess of water.

**Butter** is made by vigorously shaking (churning) cream, thus causing the globules of fat to cohere and form a compact mass. This is kneaded to remove the milk which was entangled with the fat globules of the cream, and salt is usually added as a preservative. Butter fat chiefly differs from beef and mutton fats (the basis of *margarine*) in containing a fat known as *butyrin*.

**Milk is the most perfect of all natural foods.**—Its constituents not only provide all that is necessary for complete nutrition, but are also present in very suitable proportions.

#### EXERCISES ON CHAPTER XXIV.

1. What substances are present in new milk? To which classes of food-materials do they respectively belong?
2. What is an emulsion? Give an example of a natural emulsion.
3. What would be the effect upon the specific gravity of new milk of (a) removing the cream, (b) adding water?
4. Describe the effect of adding essence of rennet to lukewarm milk.
5. What occurs when milk turns sour?
6. Of what do curds and whey respectively consist?
7. How is butter made? Compare the composition of its fat with that of beef-suet.
8. What is *margarine*? How would you distinguish it from genuine butter?
9. How is "condensed milk" prepared? In using such milk for the preparation of children's food, what precautions must be adopted?
10. Why is milk a good food for young children? What is the average composition of cows' milk? (1898.)
11. Explain the preparation of cheese. What is its value as an article of diet, and what is its average composition? (1896.)
12. Life can be sustained on milk alone longer than on any other single article of food: explain the reason for this. Why is malted food good for very young children? (1896.)
13. What is the composition of butter? Mention the common adulterations. (1899.)
14. How is bread prepared? Explain its value as a food, and contrast its general composition with that of cheese. (1897.)

## CHAPTER XXV

### THE HUMAN FRAME. DIGESTION AND DIET

#### DIGESTIVE PROCESSES

**313. Some properties of starch and sugar.**—Before this chapter is proceeded with, Expts. 214, 215, and 218 should be repeated. They show that—

(a) Starch paste will not pass through a moist membrane, such as parchment paper.

(b) Sugar solution will pass through such a membrane.

(c) Saliva will convert starch paste into sugar when the mixture is kept warm.

**314. The action of saliva prevented by acid.**—Repeat Expt. 218, but first make the mixture *faintly* acid by the addition of a little dilute hydrochloric acid. On testing it with Fehling's solution after some hours, no precipitate is obtained.

Saliva does not convert starch into sugar in the presence of free acid.

**315. The action of pepsin on white of egg.**—Into a beaker, measure 1 cubic centimetre of strong hydrochloric acid, and then add 180 c.c. of distilled water. Stir, and add a small pinch of *pepsin*, a pale yellow powder, which may be obtained from the chemists. Take a piece (about the size of a hazel nut) of fresh hard-boiled white of egg, break it up into small pieces, and add it to the solution. Put the beaker into a place where it can be kept lukewarm, and stir it at intervals.

After a few hours, or less, it will be found that the white of egg has completely dissolved.

The albumin has been changed into another proteid called *peptone*, which is soluble in water, and also passes readily through a moist membrane.

**316. Pepsin inactive in the absence of free acid.**—Repeat the last experiment, using the water and pepsin, but omitting the acid.

The white of egg remains undissolved.

**317. The presence of carbon dioxide in expired air.**—Repeat the experiment of breathing into lime-water contained in a beaker or test-tube (Fig. 111).

The lime-water is turned milky, showing that the air from the lungs contains a great deal of carbon dioxide gas.

**Organs and tissues of the human body.**—The human body consists of various parts which have special duties to perform. Thus the *skeleton* acts as a supporting framework; the *muscles* enable the body to 'make certain movements'; the *heart* pumps blood throughout the system; the *lungs* take in fresh air; the brain and other parts of the *nervous system* exert a controlling and directive influence over the activities of the rest of the body; and so on. Such parts, which perform special duties, are known as *organs*. Organs, in their turn, are built up of *tissues*, which vary, as the materials of one building may vary from those of another. Thus the heart and the muscles are composed of *muscular tissue*, the skeleton is chiefly composed of *bony tissue*, the brain consists of *nervous tissue*, etc.

**The chemical composition of the body.**—The tissues are composed of chemical compounds. These fall naturally into three groups: (1) *water*, which is driven off as steam when the tissue is dried by heating it to a temperature of 100° C. (Expt. 294); (2) *organic matter*, which burns away when the dried tissue is strongly heated in air (Expt. 295); and (3) *mineral matter*, which remains as ash when the organic matter has burnt completely away.

More than half the substance of the human body is water. The organic matter of the body consists of various compounds, chiefly built up from the chemical elements carbon, hydrogen, oxygen, and nitrogen. Some of them contain also sulphur and phosphorus. The inorganic or mineral matter is composed of compounds of the elements calcium, sulphur, phosphorus, chlorine, sodium, potassium, iron, magnesium, silicon, and fluorine. Some of these elements, however, are present in very minute quantities.

**Excretions.**—It is common knowledge that, during life, certain substances are constantly being given off from the body. The skin is always more or less moist with the perspiration which escapes from its pores. The breath which leaves the lungs contains a considerable quantity of water vapour, which condenses to visible drops of water when we breathe upon such a cold surface as a looking-glass. A large amount of water is given off in the course of the day by the kidneys. A less familiar body is *urea*, a compound containing nitrogen which

is also excreted by the kidneys. Expt. 174 proved that the air leaving the lungs is heavily loaded with carbon dioxide gas.

These waste materials, *water*, *urea*, and *carbon dioxide*, which the body gives off, are known as excretions. As water consists of oxygen and hydrogen ; urea, of carbon, hydrogen, oxygen, and nitrogen ; and carbon dioxide, of carbon and oxygen, it is evident that there is a continuous drain of the elements carbon, hydrogen, oxygen, and nitrogen from the system.

**The necessity for food.**—It is plain that, if the health of the body is to be maintained, these losses must be made good by a regular supply of new material. We obtain this fresh supply from *food*.

**Food as a fuel.**—In health, the temperature of the body never departs very widely from 37° C. or 98°·5 F. How is this temperature kept up ?

In Expt. 295 it was seen that dried beef readily catches fire and burns, giving off much light and heat. The burning is due to the combination of the carbon, hydrogen, and nitrogen of the beef with the oxygen of the air. The burning of organic matter always gives rise to heat, whatever the conditions under which it takes place. The heat may be made to do work of various kinds, as when coal is burnt in the boiler fires of steam-engines, etc. In this case the heat given off by the burning coal is partly converted into mechanical energy, which drives the engine, and, through it, does work. A precisely similar process takes place in our own bodies. *The organic matter which we take in as food is fuel.* After it has been changed into living tissue, it is gradually burnt in those parts of the body which are doing work, and the *energy or power of doing work* which the body possesses is entirely derived from the oxidation or burning of the food ; just as the energy or power of doing work which a steam-engine possesses is entirely derived from the oxidation or burning of the coal in the boiler fires.

The capabilities of the commonest food substances of acting as fuel have been determined ; and they afford a useful indication of the relative values of these foods, although, as we shall see, other circumstances have also to be taken into account. An ounce of such a proteid as lean meat or white of egg has

almost the same *fuel value* as an ounce of a carbohydrate like sugar, or starch; while the fuel value of an ounce of fat—beef-fat or butter, for example—is about  $2\frac{1}{2}$  times as great as either.

Fats are thus much more concentrated fuels than either proteids or carbohydrates.

**The maintenance of the body temperature.**—The burning or oxidation of the coal in boiler fires gives rise to heat, which is partly available, as we have seen, for doing mechanical work. Much of the heat produced, however, *remains* heat, and is not converted into mechanical energy. The same is the case with our own bodies. The energy set free by the burning of the tissues (p. 294) gives the body the power of performing all the complicated actions of which it is capable; but some five-sixths of the energy is manifested as heat, which keeps the body at a fairly uniform temperature of about  $37^{\circ}$  C.

The body maintains itself at this temperature—to a great extent automatically—by the control which the nervous system has over the circulation of the blood. Practically every part of the system is permeated by blood-vessels; and as the hot fluid circulates it keeps the body warm, somewhat as the hot-water apparatus of a school or house warms the rooms of the building.

When the rate of burning is largely increased, as it is during vigorous exercise, the supply of blood to the skin is increased, and we perspire more freely than usual. It has been explained on p. 62 that when water or any other liquid evaporates it takes up heat. The heat which the perspiration takes up as it evaporates comes from the body, which is thus cooled.

**Definition of food.**—Food may be defined as material which, when taken into the body, serves for building up new tissues. The oxidation, or slow burning, of the tissues gives rise to heat and other forms of energy. It is only in very few cases that food is directly oxidised without having been previously converted into living tissue.

**Classification of foods.**—The classification of foods which perhaps most naturally suggests itself is one into (1) animal foods, and (2) vegetable foods. This distinction is borne out by actual analysis, for vegetable foods are, as a rule, rich in one class of nutrients—the carbohydrates—which occur to only a very small extent in animal foods.

We may, however, regard food from a point of view which includes both animals and vegetables. In each case the food as obtained contains more or less material (such as bone, shells, skin, rind, bran, etc.) which cannot be used. This may be called *refuse*. The eatable part which remains may be further divided into *water* and actually nutritious substances or *nutrients*. The nutrients in their turn consist of *proteids*, *carbohydrates*, *fats*, and *mineral matters*. The classification may be exhibited in a tabular form as follows :—

<b>FOOD</b>	REFUSE ; <i>e.g.</i> , bone, shells, rind, bran, etc.						
	EATABLE PART	<table><tr><td><b>Water</b></td><td><i>Proteids</i> ; <i>e.g.</i>, albumin, casein, myosin, gluten. Consist of C, H, O, N, [S], [P].</td></tr><tr><td rowspan="3"><b>Nutrients</b></td><td><i>Fats</i> ; <i>e.g.</i>, butter fat, meat fats, oil of seeds, etc. Consist of C, H, O.</td></tr><tr><td><i>Carbohydrates</i> ; <i>e.g.</i>, starch, sugars, etc. Consist of C, H, O.</td></tr><tr><td><i>Mineral matters</i> ; <i>e.g.</i>, chlorides, sulphates, and phosphates of sodium, potassium, calcium, and magnesium.</td></tr></table>	<b>Water</b>	<i>Proteids</i> ; <i>e.g.</i> , albumin, casein, myosin, gluten. Consist of C, H, O, N, [S], [P].	<b>Nutrients</b>	<i>Fats</i> ; <i>e.g.</i> , butter fat, meat fats, oil of seeds, etc. Consist of C, H, O.	<i>Carbohydrates</i> ; <i>e.g.</i> , starch, sugars, etc. Consist of C, H, O.
<b>Water</b>	<i>Proteids</i> ; <i>e.g.</i> , albumin, casein, myosin, gluten. Consist of C, H, O, N, [S], [P].						
<b>Nutrients</b>	<i>Fats</i> ; <i>e.g.</i> , butter fat, meat fats, oil of seeds, etc. Consist of C, H, O.						
	<i>Carbohydrates</i> ; <i>e.g.</i> , starch, sugars, etc. Consist of C, H, O.						
	<i>Mineral matters</i> ; <i>e.g.</i> , chlorides, sulphates, and phosphates of sodium, potassium, calcium, and magnesium.						

**Air as a food.**—It was shown in Chapter X. that the atmosphere consists principally of a mixture of nitrogen and oxygen gases. Although nitrogen and oxygen are so necessary for building up new tissues, air can be regarded as a food only by giving the very widest—and indeed a scarcely accurate—interpretation to the term. The *nitrogen* of air is absolutely useless as food. It simply serves to dilute the oxygen, and passes in and out of the lungs unchanged. The *oxygen* of the air is a food to the following very limited extent : it combines with the elements of the tissues when they burn in the body, and so takes a material part in the liberation of energy which accompanies the burning. How does the oxygen reach the tissues ?

**Respiration.**—Respiration or breathing is carried out in the lungs (Fig. 112). The lungs have a spongy texture (Fig. 167) : consisting largely of tiny air-cells, which are separated from each other by very thin partitions. In these partitions are very small channels, called *capillaries*, through which the blood is continually flowing. The blood in the capillaries of the lungs is thus separated from the air in the air-cells by an extremely delicate membrane only ; so thin is this membrane that oxygen readily passes through it from the air-cells to the blood. The blood contains certain little red discs, called *red corpuscles* (*r*, *r'*, Fig. 168). The oxygen combines loosely with the colouring matter of the red corpuscles, and is consequently carried by them to all parts of the body as the blood circulates. As the blood passes through tissues needing oxygen, the oxygen is given up to them. The blood at the same time receives the waste carbon dioxide formed by the previous activity of the tissue ; and when this part of the blood next passes through the capillaries of the lungs, it gives up the carbon dioxide to the air, and receives a fresh supply of oxygen from the air.



FIG. 167.—A small piece of lung, highly magnified to show the air-cells (*b*).

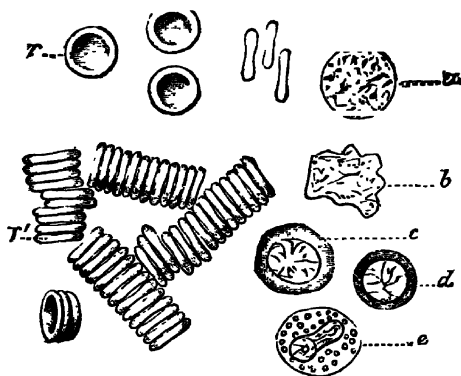


FIG. 168.—Blood corpuscles. *r*, red corpuscles seen on the flat ; *r'*, red corpuscles seen on edge, and run together into rows ; *a*, *b*, *c*, *d*, colourless corpuscles. (Very highly magnified.)

The lungs are thus not only organs of excretion, as are the skin and kidneys (p. 293), but also organs of respiration.

**Digestion.**—It has been seen that the use of food is to be



built up into the substance of living tissue, and so repair the waste of the latter. Before food can thus fulfil its purpose, it must undergo certain important changes which constitute *digestion*. The fats must be reduced to that extremely fine state of subdivision which is called an *emulsion* (p. 285); and

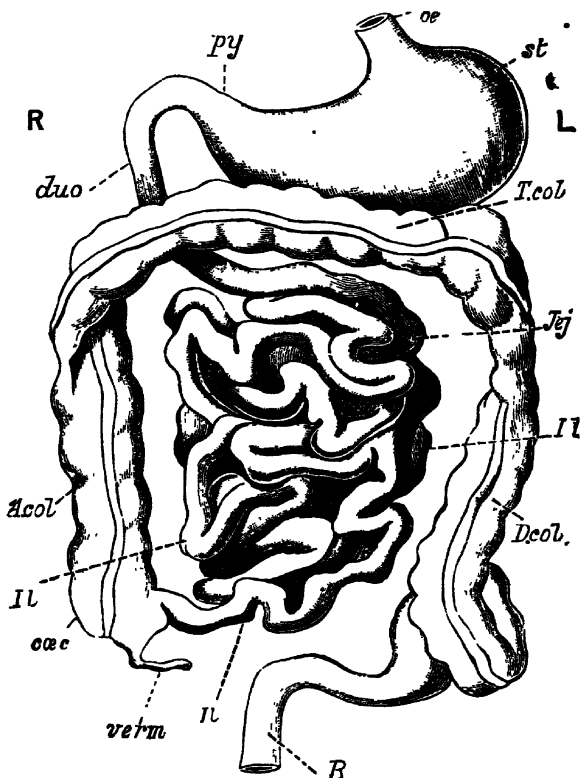


FIG. 169.—The abdominal portion of the alimentary canal.  
 R, right; L, left; æ, esophagus; st, stomach; il, small intestine; A. col, T. col, D. col, large intestine; R, rectum.

the proteids and carbohydrates must be rendered not only *soluble*, but also *diffusible*—that is, able to pass through a thin membrane. When this has been done, the digested food can be taken up by the blood, and carried to the tissues, the waste of which it is to repair.

**The alimentary canal.**—The digestion of food takes place in a tube, called the alimentary canal, which runs throughout the body. It consists of the *mouth-cavity*, the gullet or *œsophagus*, the *stomach*, and the *intestines*. The greater part of the length of the alimentary canal is shown in Fig. 169.

**The changes which food undergoes in the alimentary canal.**—We have now to consider how the constituents of such a typical meal as bread, beef, and potatoes are acted upon in their passage along the alimentary canal.

In the *mouth* the food is ground up by the teeth into fine particles. At the same time it becomes thoroughly mixed with a fluid, the *saliva*, which is poured into the mouth in great abundance during the process of chewing or *mastication*. The saliva is formed by certain organs called *salivary glands* (Fig. 170). The saliva contains a small quantity of

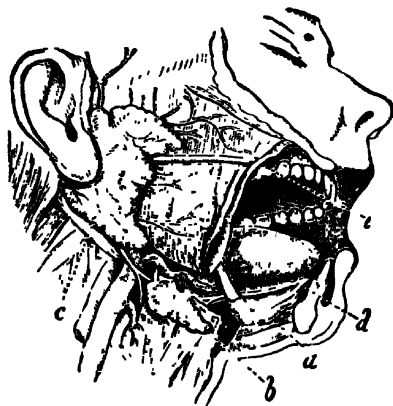


FIG. 170.—Dissection to show the salivary glands: *a*, *b*, and *c*.

*ptyalin*, a substance which has the power of changing insoluble and non-diffusible starch into a soluble and diffusible sugar (pp. 215 and 292). The food remains in the mouth for so short a time, however, that only a very small part of the starch of the bread and potatoes is converted into sugar before it is swallowed.

When the food reaches the *stomach* (Fig. 171), it is subjected to the action of another fluid, called the *gastric juice*. Gastric juice contains three important constituents; one is hydrochloric acid, which is present to the extent of about 0·2 per cent.; the other two are ferments, known as *rennet* (p. 286) and *pepsin* (Expt. 315) respectively. The hydrochloric acid arrests the action of the ptyalin of the saliva (Expt. 314) as soon as the food enters the stomach; and the action of the gastric juice is

confined to the proteids (gluten and myosin). These are converted into a soluble and diffusible proteid called *peptone* (Expt. 315).

The stomach is very muscular, and it "contracts" in such a manner that the food and the gastric juice with which it is mixed are thoroughly churned up, giving rise to a mass of the consistency of pea-soup, and known as *chyme*.

When the chyme passes through the narrow opening *d* (Fig. 171) into the first part of the small intestine, it is exposed

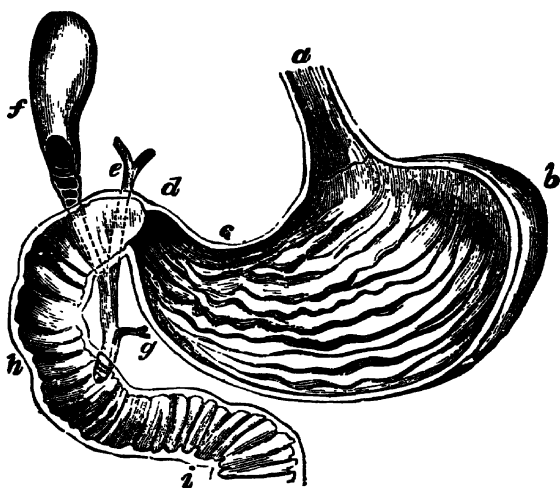


FIG. 171.—The stomach laid open. *a*, œsophagus; *b*, dilation on left end of stomach; *c*, the upper wall; *d*, the narrow opening between the stomach and the small intestine; *e*, bile duct; *f*, gall bladder; *g*, pancreatic duct, opening with bile duct into *h*; *i*, the first part of the small intestine

to the action of the secretions of two large glands, the *liver* and the *pancreas*. These secretions, known as *bile* and *pancreatic juice* respectively, enter the intestine by a common duct (Fig. 171). Both secretions contain alkalis, and these at once stop the action of the gastric juice, which can only act in the presence of free acid (Expt. 316). The digestive action of the bile is chiefly to break up the fats of the food into minute globules, or "emulsify" them (p. 285). The pancreatic juice has the same effect upon fats, but possesses in addition three ferments,

which enable it (1) to turn starch into sugar (as ptyalin does); (2) to change indiffusible proteids into peptone (as pepsin does); and (3) so to decompose fats that with the alkaline matters present they can form soaps (Chapter XXII.).

The fate of the various food-stuffs may be summarised as follows :—

*Proteids* are rendered soluble and diffusible by the action of (1) the gastric juice in the stomach; (2) the pancreatic juice in the small intestine.

*Starch* is converted into soluble and diffusible sugar (1) to a small extent by the ptyalin of saliva in the mouth; (2) by the pancreatic juice in the small intestine.

*Fats* are emulsified, and (to a very small extent) changed into soaps or saponified, in the small intestine by the pancreatic juice and bile.

**Absorption.**—When the carbohydrates and proteids have been rendered soluble and diffusible, they make their way through the very thin membrane which divides the cavity of the alimentary canal from the cavities of the minute blood-vessels permeating its walls. They then enter the blood, and are carried by its circulation to the various tissues of the body, to make good the waste which these have undergone. Absorption takes place to a small extent through the walls of the stomach, but by far the greater part is carried on in the small intestine. This applies also to the salts.

The fats are also absorbed in the small intestine, but reach the blood-stream by a somewhat roundabout path.

In the large intestine (Fig. 169) the useful material which still remains unabsorbed is taken up, together with a very large quantity of water. Such matter as is now left consists chiefly (when only so much food as is necessary is eaten) of indigestible substances. In its passage through the large intestine it gradually acquires the character of *fæces*, and is ultimately discharged by the rectum (Fig. 169).

**The daily food supply.**—An average healthy man every day gives off about  $\frac{3}{4}$  oz. of nitrogen, and about  $9\frac{1}{2}$  ozs. of carbon in the various excretions of his body. We have therefore to consider how these losses may best be replaced in his daily food. It has more than once been pointed out (pp. 210 and 274) that proteids form the only available source of nitrogen. Proteids

correspond closely to one another in composition; which may be considered to be roughly as follows :—

Carbon . . . . .	53 per cent.
Oxygen . . . . .	22 „
<b>Nitrogen</b> . . . . .	<b>16 „</b>
Hydrogen . . . . .	7 „
Sulphur . . . . .	2 „
	<hr/>
	100

As proteids form the only source from which the body can obtain its nitrogen, our average man, who throws off  $\frac{3}{4}$  oz. of nitrogen per day, must evidently consume *at least*  $\frac{3}{4} \times \frac{100}{16} = 4.7$  ozs. of proteid matter per day in order to make good this loss.

This quantity (4.7 ozs.) of proteids would thus supply the necessary amount of nitrogen, but would contain only  $\frac{4.7 \times 53}{100} = 2.5$  ozs. of carbon ; whereas the body requires some 9.5 ozs. of this latter element. How is the extra 7 ozs. of carbon a day to be supplied ? The full amount (9.5 ozs.) of carbon might (as proteids contain 53 per cent. of carbon) be obtained from  $\frac{9.5 \times 100}{53} = 18$  ozs. of proteid.

We see, therefore, that from a daily allowance of 18 ozs. of proteid matter, with water, mineral salts, and fresh air in addition, all the food necessary to support life may be obtained.

There are, however, grave objections to such a diet, for 18 ozs. of proteid contain  $\frac{18 \times 16}{100} = 2.9$  ozs. of nitrogen ; and we have seen that only  $\frac{3}{4}$  oz. of nitrogen is necessary. The surplus nitrogen is not only useless, but actually injurious, on account of the great strain which it throws upon the digestive organs and the kidneys.

It is much more economical, as regards both health and cost, to obtain the extra 7 ozs. of carbon from *non-nitrogenous foods* : fats and carbohydrates. Both consist of carbon, hydrogen, and oxygen ; but whereas fats contain roughly about 80 per cent. of carbon, carbohydrates contain about 40 per cent. only.

The 7 ozs. of carbon which we have still to provide may

therefore be obtained from either  $\frac{7 \times 100}{80} = 8\frac{3}{4}$  ozs. of fats, or  $\frac{7 \times 100}{40} = 17\frac{1}{2}$  ozs. of carbohydrates. Either would suffice, but a mixture of the two is generally found to be more acceptable to the palate. Such a mixture might, for example, consist of 2 ozs. of fat and 14 ozs. of carbohydrates.

**The importance of a mixed diet.**—It is plain, then, that although proteids alone will provide all the organic matter *necessary* for the maintenance of life, an ideal diet will consist of proteids, fats, carbohydrates, and salts in such proportions that the necessary elements are provided (1) without any undue strain being thrown upon any part of the system; and (2) in an attractive and palatable form.

**A typical meal.**—These conditions may be fulfilled in a very large number of ways, which are well exemplified in the everyday meal of beef, potatoes, and bread. The *proteids* are supplied by the myosin in the lean of the beef, and by a constituent of the bread which is known as gluten (p. 209). The rest of the beef provides the *fat*; and the potatoes, with the larger part of the bread, consist chiefly of *carbohydrates*. The *mineral matters*, or salts, are taken partly as condiments (*e.g.*, common salt), and partly unconsciously with the beef, bread, etc., of which they are constituents. The presence of mineral matter in beef was seen in Expt. 295. The *water*, whether adulterated or not, which is drunk during the course of the day, accounts almost wholly for the quantity which is normally given off from the body, although a small proportion is formed by the oxidation (p. 144) of the hydrogen occurring in all proteids, fats, and carbohydrates.

It has been found that a person can extract a larger proportion of nutrients from a mixed diet than he can from a diet which consists of a single food-material, although, judged from composition alone, the latter may appear quite as nutritious.

**Interchangeable foods.**—The selection of foods to form mixed diets is rendered easier by the fact that certain foods appear to be largely interchangeable. Fats and carbohydrates, for example, seem to serve the same purpose in the body, although the fat yields  $2\frac{1}{4}$  times the energy supplied by the same weight of carbohydrate. Among the carbohydrates, again,

sugar may usefully replace starch in certain circumstances where the latter cannot be digested with sufficient ease or rapidity. Infants, who are yet unable to digest starch, obtain their necessary carbohydrate from milk-sugar (p. 287). An ordinary lump of table sugar has about the same food-value as an ounce of potato.

The great similarity between meat and peas, beans, etc., as regards the amount of proteids present, has already been referred to (p. 276). These seeds may, therefore, to a great extent be eaten in the place of meat.

Two and a half quarts of skim-milk have almost exactly the same food-value, and contain nearly the same amount of proteid, as is possessed by one pound of steak.

**The different needs of the body.**—It seems at first sight that there are two distinct results of the activities of the body : (1) the wasting away of its tissues, and (2) the liberation of heat and other forms of energy. Adopting this view, it would appear that food is taken for two distinct purposes : (1) to replenish the parts of the body which have been used up, and (2) to provide a new supply of fuel. It was formerly thought that proteids were especially concerned with the formation of new tissue, and that the duty of fats and carbohydrates was to act as fuel for keeping up the supply of heat. Thus proteids were called *tissue-formers*, and fats and carbohydrates *heat-producers*.

This view is now known to be an erroneous one. The production of heat is not a separate phenomenon, but is a necessary result of the oxidation, and consequent wasting, of the substance of the muscles and other tissues. Moreover, fats and carbohydrates, as well as proteids, play an important part in building up new tissue ; and the distinction between nitrogenous and non-nitrogenous food-stuffs is lost so soon as they have become built up into the substance of the tissues. Both are tissue-formers, and, since heat is produced by the oxidation of the tissues, both are heat-producers.

It is, however, admitted that certain classes of foods are more suitable than others for certain conditions of life. Fats are especially useful in cold countries. Sugar in small doses is of great assistance at times of unusually great muscular labour. People of sedentary habits require food containing a large proportion of proteids.

**Comparative digestibility.**—In selecting foods it is necessary to attach great importance to the ease or otherwise with which the body can bring their ingredients into a soluble and diffusible form. In other words the *digestibility* of the food, as well as its composition, must be considered. Comparatively little precise information upon this point has yet been obtained, but a few general conclusions may be mentioned.

Animal foods are, as a rule, more easily digested than those of vegetable origin, and this applies particularly to proteids. A man in ordinary health can generally digest practically all the proteid matter of meat and fish, whereas 25 per cent. of the proteids of peas and beans may remain undigested, and thus be wasted. This may be largely explained by the fact that the vegetable food-stuffs are usually enclosed in hard cell-walls, upon which the digestive juices cannot act (Figs. 162 to 164). Carbohydrates, such as sugar and starch, are easily digested when the walls in which they are enclosed are ruptured by cooking. Lean meats appear to be more easily digested than fatty meats, and beef is slightly more digestible than lamb or fish. Fresh fish is more thoroughly digested than salt fish.

**Condiments and stimulants.**—It is customary to add to food various condiments and sauces. With the exception of salt, which is a necessity, these must be regarded as luxuries of very minor importance to healthy persons: digestion taking place as well in their absence. With invalids, however, the case is different, and the use of condiments is often a powerful aid to digestion by stimulating the flow of the digestive juices.

Tea and coffee, and meat extract (p. 278), are stimulants—not foods. Taken in moderation, they are without injurious effects on digestion.

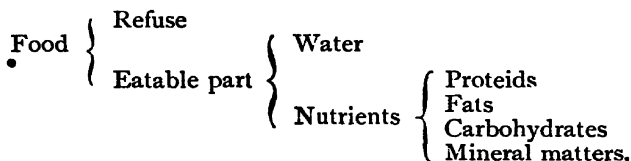
**Fruits and vegetables.**—Potatoes, turnips, carrots, beet-roots, bananas, etc., are foods in the true sense of the word, but cabbage, lettuce, tomatoes, apples, oranges, strawberries, etc., derive their value chiefly from the *mineral salts* and *organic acids* (such as malic, citric, and tartaric acids) which they contain. Some of these compounds have considerable medicinal value.



## SUMMARY.

The human body consists of (1) *water* (oxygen and hydrogen); (2) *organic matter* (carbon, hydrogen, oxygen, nitrogen, sulphur, and phosphorus); (3) *mineral matter* (calcium, sulphur, phosphorus, chlorine, sodium, potassium, iron, magnesium, etc.).

**Food** is eaten to supply the elements which the body excretes as the waste products of its activity. It also acts as a *fuel* by building up tissues which are afterwards oxidised.



**Digestion** consists in rendering food soluble and diffusible.

*Proteids* are digested (1) in the stomach by the gastric juice; (2) in the small intestine by the pancreatic juice.

*Fats* are digested in the small intestine by the pancreatic juice and bile.

*Starch* is digested (1) partly in the mouth by the ptyalin of the saliva; (2) in the small intestine by the pancreatic juice.

The *absorption* of digested foods into the blood takes place chiefly in the small intestine.

Although *proteids* supply all the necessary elements of food, a *mixed diet* is most economical, and is also very advisable from considerations of health.

All foods are *tissue-formers*, and therefore all are *heat-producers*.

*Animal foods* are, as a rule, more easily digested than *vegetable foods*.

## EXERCISES ON CHAPTER XXV.

1. What is the action of saliva on starch? How can you prove that this action really takes place?

2. What is the action of pepsin on proteids in the presence of a very dilute solution of hydrochloric acid? Where does this action occur in the body?

3. About how much nitrogen and how much carbon does an average man give off in a day? In what form are they respectively excreted?

4. What are the objections to a diet consisting wholly of lean meat?

5. Draw up a table showing the food materials contained in a meal of bread, meat, and potatoes.

6. What do you mean by digestion? Describe the digestion of a meal of bread and milk.

7. How is the heat of the body maintained?

8. Define : organ, tissue, food, and ferment. Give examples.

9. In what sense is food a fuel?

10. What is a standard diet? Give two examples of a simple meal containing all the constituents of food in about the right proportion. (1897.)

11. Give a classification of food substances, with examples, and explain their respective uses. (1897.)

12. What is the composition of beef? In what respects would a diet of 1 lb. of bread and 12 oz. of meat per day be deficient?

What amount of proximate alimentary principles should be supplied in a diet for ordinary work? (1898.)

13. What are the functions of nitrogenous food? Mention the chief nitrogenous principles to be met with in food, and their relative values as nutrients. How much nitrogen is required by an adult at ordinary work? (1894.)

14. What is the general composition of fat, and how does it differ from that of sugar? What is meant by the digestion of fat; when and how does it take place. (1895.)

15. What are the essential constituents of a diet for ordinary work? Explain the functions fulfilled by each constituent, and how far one constituent may replace another. (1899.)

16. What are the chief mineral salts contained in food? Explain their general uses. (1897.)

## CHAPTER XXVI

### MICRO-ORGANISMS AND THEIR WORK

REFERENCE has been made in earlier chapters to minute forms of life present in the air, in milk, etc., which grow and multiply with great rapidity when they gain access to liquids which afford suitable food materials. Yeast (Fig. 153), the vinegar organism (Fig. 158), and the lactic acid bacterium (Fig. 166) thrive in weak solutions of sugar, in weak alcohol, and in milk respectively.

By the use of specially prepared food substances or *nutrient media*, many other bacteria may be encouraged to grow and form colonies. Such a food material is meat-broth.

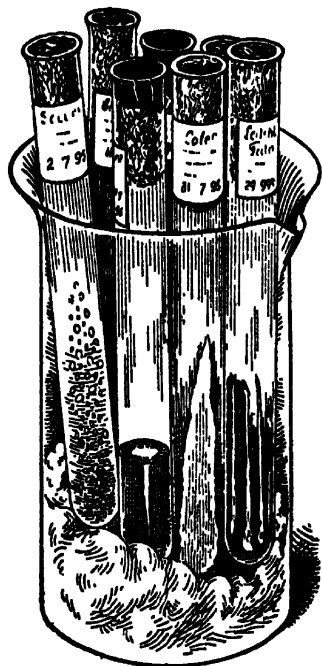


FIG. 172.—A beaker containing several culture tubes.

### THE CULTIVATION OF BACTERIA

**318. Nutrient Gelatine.**—Obtain a dozen test-tubes of nutrient gelatine (Fig. 172) from a dealer in bacteriological apparatus. This substance is meat-broth, or *bouillon*, to which about

10 per cent. of the best gelatine has been added, in order to keep it solid at the ordinary temperature of the room. Put one of the tubes, without removing the plug of cotton-wool with which it is closed, into water, and gently heat the water by means of a small Bunsen flame. Determine the temperature at which the gelatine melts.

**319. Petri's dishes.**—Obtain a few Petri's dishes. These are round glass dishes (Fig. 173), about three inches in diameter and half an inch deep, with glass lids fitted over them.

Wash the dishes and lids, dry them, *put on the lids*, and heat them for an hour in the air-oven at a temperature of  $170^{\circ}\text{C}$ . All forms of life which might have been present on the dishes are thus destroyed, and the dishes are therefore said to have been *sterilised*.

So long as the lids are left on, the insides of the dishes will remain sterile.



FIG. 173.—Petri's dish.

**320. Presence of bacteria in air.**—

Melt the gelatine of one of the tubes, take out the cotton-wool, and pour the gelatine into a Petri's dish. Leave off the lid, and let the dish remain with the gelatine—which soon solidifies—thus exposed to the air for an hour. Then put on the lid, and keep the dish for two or three days at a temperature of from  $18^{\circ}\text{C}$ . to  $20^{\circ}\text{C}$ .

A number of specks will appear in the gelatine. These are *colonies* of bacteria. Each one has been formed by the multiplication of a single cell which dropped in from the air. While the colonies may be seen by the naked eye, the individual bacteria of which they are composed can only be distinguished by using very high powers of the microscope.

Fig. 174 shows the appearance of colonies of bacteria when moderately magnified.

**321. Bacteria unable to pass through plugs of cotton-wool.**—The fact that the gelatine in the tubes remains clear, shows that bacteria are prevented by the plugs of cotton-wool from getting into the tubes from the air.

**322. Presence of bacteria in water.**—Melt the gelatine in a tube, set the cotton-wool plug on fire, and immediately put out the flame by pushing the wool further into the tube, using for this purpose a pair of forceps (tweezers) which have been sterilised immediately before by passing them through the Bunsen flame. Now take out the plug by means of the forceps, and quickly add about one cubic centimetre of tap-water to the gelatine. Fire the cotton-wool again, and push it, still burning, into the upper part of the tube, when the flame, of course, goes out. This treatment keeps the mouth of the tube sterile.

*Gently* shake the tube, to mix the water and gelatine, avoiding the formation of froth, and not allowing the liquid to wet the cotton-wool plug.

Take out the plug, and quickly pour the contents of the tube into a sterile Petri's dish, raising the lid slightly at one side to allow the liquid to be added.

Close the lid at once, and keep the dish for a couple of days at a temperature from  $18^{\circ}\text{C}$ . to  $20^{\circ}\text{C}$ .

Colonies make their appearance in the gelatine. These are formed by the multiplication of the bacteria which were present in the tap-water.

**323. Bacteria absent from boiled water.**—Half fill an ordinary flask with similar tap-water, and boil it gently for an hour. Whilst it is still boiling, thrust a burning plug of cotton-wool into the neck of the flask, and at once remove the burner from beneath the flask.

Allow the water to cool, and then, by means of a sterilised pipette, add about one cubic centimetre of it to a tube of gelatine (as in Expt.

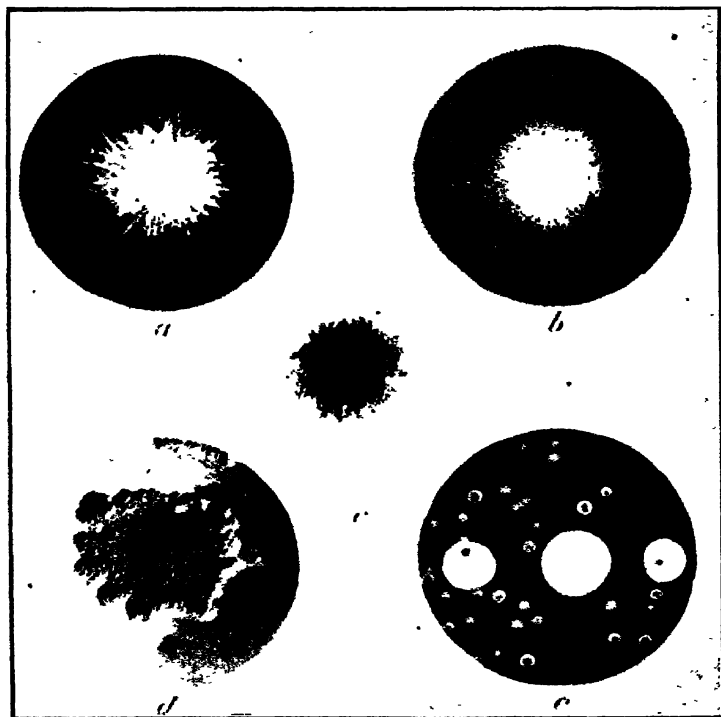


FIG. 174.—Colonies of Bacteria. *a*, *b*, *c*, and *d* each show one colony; *e* shows three large colonies and several small ones. (Reproduced from photographs taken by Mr. J. E. Barnard.)

322), and keep the mixture in a Petri's dish at about 20° C. for two days.

The gelatine will probably remain clear.

**324. Bacteria present in soil.**—Add a few grains of fine garden soil to a tube of liquefied gelatine, mix, and keep in a Petri's dish as before

Colonies of bacteria form in the gelatine.

**325. Bacteria absent from ignited soil.**—Ignite a little of the same

soil in a porcelain crucible for fifteen minutes (Fig. 175). Cool the crucible with the lid on, and then test the soil for bacteria as in the last experiment.

No colonies of bacteria are formed.

**326. Bacteria present in milk.**—Add a drop of milk to a tube of liquefied gelatine, observing the usual precautions, and keep this in a Petri's dish for two days at about 20° C.

Colonies of bacteria are formed.

**327. Bacteria absent from boiled milk.**—Raise the temperature of some milk to the boiling point, plug the neck of the flask with burning cotton-wool, and cool. Test a drop with gelatine as in the last experiment.

The gelatine probably remains free from colonies.

**Bacteria.**—The air we breathe, the water we drink, and the soil beneath our feet, are swarming with myriads of tiny living things, which exert an enormous influence upon the welfare of the whole human race. To the services of some of them we owe not only many luxuries, but even the possibility of life; while others are responsible for most, if not all, of the infectious diseases to which we are subject.

The possessors of these boundless powers for good and evil are extremely minute *plants*: each individual consisting of one cell. They vary greatly in size, but not even the largest can be seen with the naked eye; while the smallest known require magnifying by the microscope to about a thousand times their real length, breadth, and thickness to render them visible. The bacteria shown in Figs. 176 and 177 have been magnified to 1500 times their real size.

These little plants are called *bacteria*, or microbes. When a bacterium is full-grown, it multiplies by breaking up into two or more parts, each of which becomes a complete bacterium. As some bacteria become adult in half an hour or less, it is plain that, in favourable circumstances, reproduction takes place with enormous rapidity. As a matter of fact, one individual can thus give rise to about 17,000,000 in twenty-four hours.

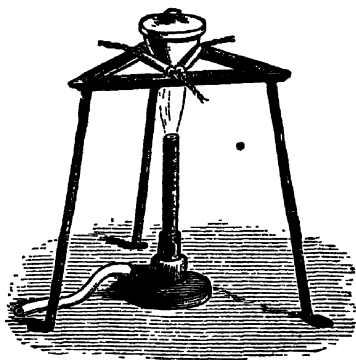


FIG. 175.—Soil may be sterilised by heating it with the Bunsen flame.

When the supply of food runs short, or the surroundings become in other respects unfavourable, many bacteria form bodies called *spores* (Fig. 176), which are very difficult to kill. They remain patiently in this stage until the hard times are over, and then, with unimpaired vigour, resume their ordinary mode of life.

**Putrefaction and decay.**—It is well known that when a piece of meat is left freely exposed to the air it soon becomes

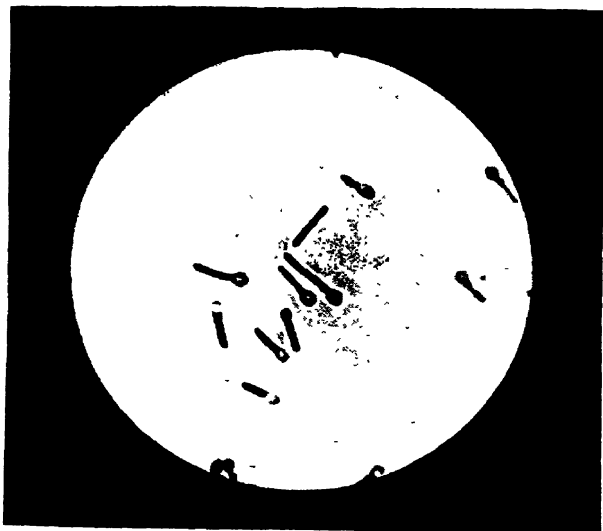


FIG. 176.—Bacteria of Tetanus (lockjaw). ( $\times 1500$ .) The rounded bodies at the ends of the bacteria are spores; they can withstand boiling for five minutes. These bacteria of tetanus occur abundantly in garden soil. (Reproduced from a photograph taken by Mr. J. E. Barnard.)

offensive. The change is due to the activities of myriads of microbes, which, by using the meat as food, break it up, particle by particle, into simple and harmless compounds. Some of these compounds are gases which pass into the air; others are at last washed into the soil. The change is largely one of oxidation, and is very similar to what would be produced by burning the meat.

The services which bacteria perform in decomposing and

thus clearing away accumulations of rubbish can scarcely be over-estimated.

That the putrefaction of organic matter is really due to minute and air-borne forms of life was conclusively proved by Pasteur and Tyndall about the middle of the nineteenth century. It was shown that if well-boiled broth is kept in vessels from which the air is either wholly excluded or—by plugging the mouth of the vessel with cotton-wool, for example—so admitted

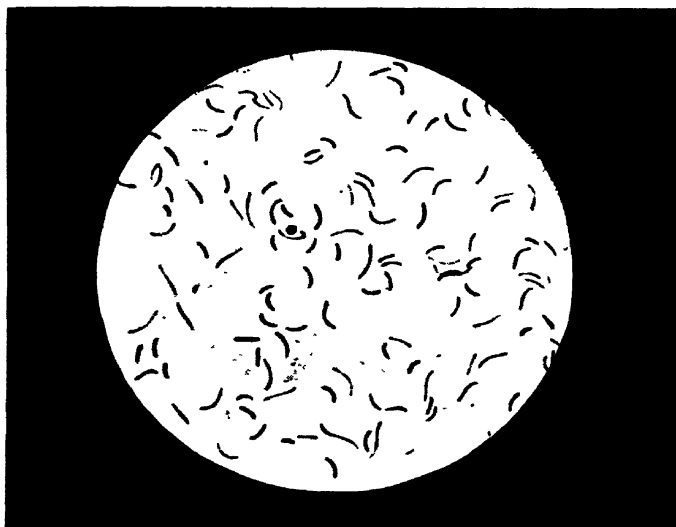


FIG. 177.—Bacteria of Cholera. ( $\times 1500$ .) (Reproduced from a photograph taken by Mr. J. E. Barnard.)

that all floating particles are arrested, no putrefaction occurs, and the broth remains sweet for an indefinite time.

**Sterilisation.**—During the boiling of the broth all forms of life which were originally present in it are destroyed. Broth so treated is said to be *sterilised*.

**The presence of bacteria** in water, soil, and milk and other articles of food may be easily proved by means of *nutrient gelatine*, which may be obtained from dealers in bacteriological apparatus. It is usually supplied in test-tubes, which are sealed from dust (and therefore from bacteria) by plugs of cotton-wool



(Fig. 172). The substance consists of broth, which has been mixed with about 10 per cent. of the best gelatine and sterilised. The nutrient gelatine is melted, mixed with a little of the substance to be examined for bacteria, and then poured into a "Petri's dish" (Fig. 173) which has been previously sterilised by heating to  $170^{\circ}$  C. for an hour. The dish—covered with its lid—is kept for two days at a temperature of from  $18^{\circ}$  C. to  $20^{\circ}$  C. At the end of that time the presence of bacteria is shown by the appearance of specks in the gelatine. These specks consist of *colonies* of bacteria. Several such colonies, as seen under low magnifying powers, are shown in Fig. 174.

**Harmless and injurious bacteria.**—Such micro-organisms as *yeast*, the *vinegar organism*, and the *lactic acid bacterium*, which have been referred to in previous chapters, are harmless; but others set up changes in food which result in the production of deadly poisons called *ptomaines*. These compounds are found from time to time in meat, fish, cheese, ice-cream, etc. Still other bacteria are the cause of the infectious diseases, and, by gaining access to the bodies of healthy persons, may be the means of spreading the diseases.

**Necessity for sterilising food.**—Injurious bacteria are most commonly distributed in water, milk, and other articles of food, and too much care cannot be taken to avoid using food which is liable to such contamination. The most satisfactory means of destroying bacteria in food and water is to heat these latter to the temperature of boiling water. Suspected water ought always to be boiled before being drunk. The "flat" taste produced by boiling may be removed by afterwards pouring the water backwards and forwards from one vessel to another, so as to mix it thoroughly with air. *Milk* has so often been proved to have conveyed disease, that it is always safest to raise its temperature to the boiling point before it is used for food.

**Tinned foods.**—Properly prepared tinned foods are, from the method of their preparation, sterile. The tins of cooked food, with one small aperture, are surrounded by a solution of zinc chloride, which boils at a higher temperature than water. The tins are heated in this, and when all the air has been expelled by steam, are sealed by a drop of solder, and cooled. If the heating has been sufficient, the contents will remain free from

decomposition for any length of time. Small tins are more likely than large ones to have been thoroughly heated through, and on that account are to be preferred. Tins which have bulged out should be rejected, as this indicates decomposition and the formation of gas in the contents.

When a tin is opened, the contents should be at once removed to a glass, porcelain, or earthenware vessel. Tinned fish should be eaten at once.

**Preservation of food.**—Food may be preserved for an indefinite period if it be kept *frozen*, since cold destroys some bacteria, and suspends the activities of the rest. Meat is, for example, brought from Australia and New Zealand in ships fitted with refrigerating chambers.

*Moisture* is necessary for the active life of bacteria, and therefore *dried* meats and fish may be kept for a long time without undergoing decomposition.

*Salt* and *saltpetre* and even *sugar* are widely used in "curing" meat and fish. The action of a strong solution of sugar in preventing fermentation is also taken advantage of in preserving fruits, etc. Borax, boracic acid, salicylic acid, and other substances are also used for the preservation of various articles of food. They are injurious if present in more than small quantities.

**Necessity for cleanliness.**—Disease and dirt are closely connected. Accumulations of refuse, grease, etc., form the most favourable surroundings for injurious bacteria; and bacteria cannot be more effectually exterminated than by scrupulous attention to the cleanliness of the person and the surroundings. *Cooking-utensils*, ovens and stoves, etc., especially, should be carefully cleaned as soon as they have been used. The great value of water, soda, and soap for removing dirt and grease has already (p. 267) been referred to. It is worthy of notice that *light* has a very injurious effect upon bacteria. As much light as possible ought to be admitted to all the rooms of a house.

**Prevention of contagion.**—The transmission of disease may be largely avoided by taking care that the patient is properly isolated, and that all clothing, utensils, etc., which have been used in the sick-room are thoroughly disinfected.

So far as possible, all clothing should be well boiled. Garments, bedding, etc., which would be spoiled by this treatment,

should be heated in the sterilisers which most municipal authorities possess. In a common form, the garment is subjected to the action of steam under pressure. This permeates the pores of the cloth, and destroys all germs without charring or otherwise injuring the fabric.

Utensils ought to be disinfected by means of a solution of corrosive sublimate, carbolic acid, Condy's fluid, or other germicide.

Rooms in which serious infectious diseases have occurred ought to be fumigated by the vapour of formic aldehyde, chlorine gas, or the fumes of burning sulphur. Of these, perhaps formic aldehyde, which has come into use comparatively recently, is the most effective and the least objectionable.

Disinfection and fumigation are usually carried out by qualified officers of the local sanitary authority.

#### SUMMARY.

**Bacteria** are minute, one-celled *plants*, which are found almost everywhere. They are destroyed by heat.

Some bacteria cause *putrefaction* of organic matter. Poisons called *ptomaines* are often produced during the process. Food which has begun to decompose should not be used for food.

Infectious diseases are due to certain bacteria which gain access to the body.

**Sterilisation** of food is the destruction of the bacteria present in it. Food may be sterilised by boiling. Small tins of tinned meat, fish, etc., are more likely to be sterile than large ones.

**Food is preserved** (1) by being sealed from air after sterilisation by heat; (2) by freezing; (3) by the addition of substances which prevent the growth of bacteria.

The propagation of bacteria and the spread of disease are best prevented by *cleanliness* and *light*.

#### EXERCISES ON CHAPTER XXVI.

1. What are bacteria? Where do they occur, and how are they propagated?

2. How would you prove the presence of living particles in (a) air, (b) milk?

3. What is the cause of putrefaction? What changes occur in the composition of a piece of meat when it putrefies?

4. What are ptomaines? In what articles of food have they been known to occur?

5. Why should food be sterilised? How would you sterilise (a) water, (b) milk?
6. Describe the preparation of tinned foods. What precautions should be observed in their use?
7. Why do frozen meat and salt fish keep longer than fresh meat and fish?
8. Why is sugar used for preserving fruit?
9. Why is it necessary to clean cooking utensils immediately after using? How should the cleaning be done?
10. How are clothing and rooms disinfected?

# TYPICAL EXAMINATION PAPERS.

## BOARD OF EDUCATION. .

### ELEMENTARY SCIENCE OF COMMON LIFE.

#### INSTRUCTIONS.

*You are permitted to answer only eight questions, of which six must be taken from division A and two from either B or C.*

#### A.

1. How would you show that an "empty" bottle is full of air? What are the effects of heat and of pressure upon air, and by what means can alterations in the temperature and pressure of the atmosphere be ascertained?

2. If a colourless liquid were given you, by what tests would you prove it to be water?

3. Describe carefully what you see when a candle burns. What are the principal changes that take place in the candle-flame, and by what experiments could you prove that they occur?

4. If 20.5 cubic centimetres of a solution of common salt in water weigh 23.5 grams, what is the specific gravity of the solution?

5. Describe a common thermometer and state how you would proceed to test its accuracy.

6. What is limestone? What substances are obtained when limestone is (a) heated strongly in the air, (b) acted on by hydrochloric acid, and how do they differ from it in their behaviour towards water?

7. How is coke made, and of what does it chiefly consist? In what respects does a coke fire differ from a coal fire, and how is the difference explained? What does coal ash consist of?

8. Describe the principal characteristics of the two gases, carbon dioxide and carbon monoxide, and how each gas is recognised when pure.

What important difference is to be noticed in the effect of these two gases on men and animals?

9. By what tests do you distinguish an acid from an alkali?

Classify into acid and alkaline the following substances: spirit of salt, washing soda, ammonia, common soap, slaked lime, borax.

## B

10. If you were asked to classify an elementary substance as a metal or non-metal, what experiments would you make?

11. What is the cause of the tarnishing of a copper kettle or saucepan when placed on a clear fire? Why is it necessary to avoid keeping vegetable matters for any long time in copper pans?

12. Name the oils used for illuminating purposes, and describe what would happen if each in turn were boiled with a solution of caustic potash or soda.

13. How would you distinguish from one another bars of cast iron, wrought iron, and steel?

## C.

14. From what sources is sugar obtained? Give an outline of the process by which it is made.

15. What substances are produced from sugar by fermentation with yeast, and how can these be separated and identified?

16. How is common hard soap made? Explain the effect of adding some vinegar or other acid to a solution of soap.

17. How can starch be obtained from flour, and how would you show that substances obtained by a similar process from rice and potatoes are probably identical in composition with wheat starch?

## II

## A.

1. What is meant by specific gravity? Calculate the specific gravity of a liquid 30 c.c. of which weigh 27 grams.

2. By what experiments would you prove that water is composed of nothing but hydrogen and oxygen?

3. How would you obtain nitrogen from air and distinguish it from carbon dioxide?

4. What properties are common to acids? Show that these are exhibited by the volatile product obtained on heating saltpetre with sulphuric acid.

5. How are the degrees  $0^{\circ}$  and  $100^{\circ}$  on the scale of a Centigrade thermometer determined? State the corresponding temperatures on a Fahrenheit thermometer.

6. When steam is decomposed by red hot iron, how could you prove that one constituent of the steam combines with the iron?

7. What experiments would you make in order to show that atmospheric air contains a practically constant proportion of oxygen?

8. If supplied with a nearly white powder supposed to be slaked lime, what experiments would you make to determine the question? How would lime be distinguished from lime-stone?

9. Draw a diagram of a Bunsen burner and explain its action.

## B.

10. Why is iron suitable and lead unsuitable for the manufacture of saucepans?

11. Why is it that coal gas does not explode when burned in the ordinary way or in a jar turned mouth downwards? What conditions are necessary for an explosion to occur?

12. What are the following alloys composed of: silver coinage, brass, pewter, solder? Describe characters or tests by which the constituents of *one* of them may be recognised.

13. What is iron rust? Give any explanation you can of the process of rusting.

## C.

14. What is the substance that renders vinegar acid? How could you separate it from vinegar, and prove that it enters into the composition of sugar of lead?

15. What happens when a fat is boiled with caustic soda, and when an acid is added to the product? What is the cause of the curd produced when soap is used with hard water?

16. How would you prove the presence of carbon and hydrogen in such a substance as sugar, and of carbon, hydrogen, and nitrogen in cheese?

17. What experiments would be necessary to demonstrate the presence in ordinary beer of (a) carbon dioxide, (b) alcohol, and (c) sugar?

## III

## A.

1. Write down the equivalents in English weights and measures of (a) 760 millimetres, (b) 10 cubic decimetres, (c)  $2\frac{1}{2}$  kilograms.

2. A block of wood measuring  $10 \times 7 \times 4$  centimetres weighs 252 grams. What is the specific gravity of the wood?

3. How do you explain the presence of water on the inside of windows during a cold night while everything else in the room is apparently dry and warm?

4. Describe a common thermometer and the process of making the scale of degrees marked upon it.

5. Explain the production of flame in an ordinary domestic fire. Why does a coke fire produce very little flame?

6. Explain the difference between hard and soft water, and describe two methods for softening hard water.

7. A white crystalline powder is supposed to be either borax or washing soda; what tests could you apply to distinguish these salts from each other, and how would both be distinguished from common salt?

8. What is quicklime? How is it obtained? What happens when

water is added to it, and when it is exposed to the air for a long time?

9. Describe two experiments to prove that air expelled from the lungs contains the same gas as that produced by burning carbon in oxygen.

B.

10. What is an alloy, and how are such substances made? Name four alloys in common use, and the constituents of each.

11. Why is sulphur classified as a non-metal and copper as a metal?

12. Distinguish between the results obtained when coal is burned freely in the air, and when it is heated in retorts as at the gasworks.

13. Describe the effects of heating lead and copper strongly in the air. Would it make any difference if they were surrounded by coal gas during the process?

C.

14. From what sources is acetic acid obtained? Why is it called an acid?

15. How can it be proved that gluten or albumen contains carbon, hydrogen and nitrogen?

16. What products are formed by boiling fat with caustic soda, and how are they separated from each other? Describe the chief properties of each.

17. Explain the use of yeast in brewing and in bread-making.

IV.

A.

1. A gallon of water weighs 10 pounds; calculate the number of grams of water and the number of litres in a gallon.

2. How many yards are there in a kilometre, and how many metres in an English mile?

3. Describe experiments to show that solids, liquids, and gases expand when heated. If a corked bottle containing only air is made hot the cork may be driven out; why is this?

4. Why does ice form first on the surface of a pool of water in frosty weather?

5. The ground being covered with snow, and the sun shining, pieces of black and white cloth are laid on the snow; describe and explain the effect in each case.

6. What are the chief gases in the atmosphere, and how can each be obtained separately in a pure state? Describe the effect of each on a lighted candle.

7. Why is sea water so much more salt than river water? How would you make pure water from sea-water? Draw a diagram of the apparatus required.

8. How are chemical compounds distinguished from mixtures of the same materials? A mixture of 7 parts of iron filings, with 4 parts of



sulphur has the same composition as iron sulphide; illustrate your answer by reference to this case.

9. How would you show that the gas evolved from a lime kiln is the same as that which is produced by burning charcoal in air? Is it right to speak of "burning" lime?

10. How would you distinguish by tests, common salt from washing soda, slaked lime from quicklime, and slaked lime from chalk?

### B.

11. Enumerate the constituents of brass, pewter, silver money, and describe the chief properties of the separate metals.

12. Under what conditions does iron rust, and how can rusting be prevented?

13. How would you test a sample of paraffin oil to see if it was fit for use in a lamp?

14. If an escape of coal gas is suspected in a house, how would you proceed to find out the point of escape, and to stop it?

### C.

15. Explain the difference in composition between a fat and a carbohydrate, such as common sugar. What experiments could be made to illustrate the difference?

16. Make a drawing of the appearance of any one kind of starch under the microscope. What happens if starch is mixed with (a) cold water, (b) boiling water, (c) infusion of malt?

17. How could you prove the presence of alcohol in beer, and how could you find out, approximately, the amount present?

18. How could you test for nitrogen in an organic substance? Mention three or four important articles of food which contain nitrogen, and would, therefore, answer to this test.

# UNIVERSITY OF LONDON.

## JUNIOR SCHOOL EXAMINATION.

*[Six questions only are to be attempted.]*

1. Soap lathers easily in some kinds of water and in others with difficulty only. Explain this difference as fully as you can. How could water in which soap will not lather be changed into water in which soap lathers easily?

2. What experiments would you perform to find out if the air in which a candle burns undergoes any change?

3. Describe the properties of white of egg and the change which it undergoes on heating. To what class of foods does this substance belong? Give a reason for your answer.

4. You are directed to superintend the boiling of a leg of mutton for dinner. Explain how you would proceed, and give reasons for the precautions you would adopt.

5. Bread has been called "the Staff of Life." Account as well as you can for this name, explaining what it contains and why it is valuable as a food.

6. Name six articles of food which contain a great deal of starch. Why are these foods cooked before being eaten? What changes does starch undergo in the human body?

7. Describe a satisfactory way of ventilating a sitting-room and explain why ventilation is necessary.

8. What part, if any, of its food does a green plant obtain from the air? In what form and in what circumstances is such food taken in?

9. Write a short essay on "Putrefaction."

10. Explain as fully as you can the importance of Oxygen to animals and plants.

## COLLEGE OF PRECEPTORS.

### CERTIFICATE EXAMINATION. (THIRD CLASS.)

*[Not more than Six questions are to be attempted.]*

1. What do you understand by a unit? What would you consider suitable units for the measurement of the weight of a small glass bead, the volume of a lead shot, and the area of a grass lawn respectively?

2. Give an account of the obvious properties of common clay, and point out the properties which make clay of commercial value.

3. On a sheet of cardboard a figure bounded by an irregular curved line is drawn. How would you determine the exact area of the figure enclosed by the line?
4. Upon what conditions does the floating or sinking of a body in a liquid depend? A substance of density 1.24 sinks very slowly in a certain liquid, whilst a substance of density 1.22 is observed just to float on the same liquid. What inference can be drawn from these observations as to the density of the liquid?
5. How would you arrange an experiment to show that heat is absorbed when ice melts?
6. Sketch an apparatus such as you would use for preparing pure water from sea-water, and fully describe its various parts.
7. What effect is produced on the composition of air by allowing iron to rust in it? Describe experiments you would make with the object of verifying your statement.
8. Describe experiments which conclusively show that hydrogen is one of the constituents of water.

## TYPICAL QUESTIONS FROM OTHER EXAMINATION PAPERS.

### TEACHERS' CERTIFICATE EXAMINATION.

1. Calculate the volume in litres of a cube whose edge is equal in length to the width of this Question Paper. Give the weight in grammes of this volume of water.
2. A porcelain weight of one pound breaks into two unequal fragments. How could you, without using a spring balance or ordinary balance, determine the weight of each fragment?
3. Explain how you would ascertain (1) the volume and (2) the specific gravity of an apple or potato.
4. A glass bottle weighs 6 grammes and holds 25 grammes of water. Calculate the volume of the glass itself and the volume of the air displaced by the bottle when filled with water.  
[Specific gravity of glass = 2.5.]
5. A nurse cleanses a doctor's thermometer, which reads up to 105° F., in boiling-hot water. The doctor now finds that the thermometer is useless. Why is this?
6. Describe experiments which you have seen illustrating the conversion of (a) a solid into a gas; (b) a liquid into a solid; (c) a liquid into a gas. Describe what you have observed during each process. [Reference to any of the forms of water must be excluded from your answer.]
7. A saucer containing water is left to evaporate on a window-sill. Explain the atmospheric conditions which will favour or retard the disappearance of the water.

8. How would you show that air is a mixture, and not a compound, of oxygen and nitrogen? How would you obtain nitrogen from air? By what process can oxygen be obtained?

9. How would you show experimentally that one of the constituents of water is also a constituent of air?

10. Carbonic acid gas is allowed to bubble continuously through lime water. State what you may observe when this experiment is performed. On what natural appearances does such an experiment throw light?

11. What are the circumstances which make the air of a room unhealthy for breathing? How can the conditions be met and improved?

12. Describe what may be observed when starch is strongly heated in a test-tube. What is the nature of the substance left, and how would you identify it experimentally?

## NATIONAL FROEBEL UNION.

1. Why do some bodies float in water and some sink? Why do iron ships float?

Describe *any* simple way in which you can show that a body weighs less in water than in air.

2. How can you show (i) that air expands when heated; (ii) that warm air, when surrounded by colder air, rises; (iii) that air fills apparently empty vessels?

3. Explain scientifically the following line in Shelley's poem, "The Cloud":

"I [the cloud is speaking] change, but I cannot die."

4. (i) What changes take place in the water of a pond during a very gradual change of temperature from about  $45^{\circ}$  to  $28^{\circ}$  Fahrenheit? (ii) What is meant by the boiling point of water? By what conditions is the boiling point affected?

5. How can you make a simple mercury barometer? Explain how such an instrument can be used to measure the pressure of air, and can act as a weather glass.

6. How would you regulate the ventilation and temperature of a small schoolroom (a) in summer, (b) in winter?

7. Which do you consider the best for the heating of a schoolroom—hot-air pipes, open fire-places, or closed stoves? Give your reasons.

8. What are the chief chemical constituents of the body, and in what form are they usually supplied in children's diet during the first ten years of life?

9. Mention the food principles necessary to good bodily nutrition. Which of these are the more costly? In what forms of food can they be supplied economically to children of six years old?

10. What are the constituents of milk? Why is milk food good for young children?

11. Enumerate the chief advantages obtained by the cooking of food. Which forms of cookery are the most wholesome?

12. Briefly describe the changes which take place in a piece of meat while it is undergoing digestion.

13. What is digestion? What secretions take part in the process? What are nitrogenous and non-nitrogenous foodstuffs, and what part do they play in keeping the body in good health?

14. What are the Salivary Glands? Where are they situated? What effect has Saliva on the food?

## ROYAL SANITARY INSTITUTE.

1. What do you understand by the relative humidity of the atmosphere, and how is it measured? Describe in detail any hygrometer with which you are acquainted, and the method of using it.

2. What are the chief causes of air pollution in large towns? What measures are available for minimising the evil?

3. What do you understand by the terms "radiation," "conduction," and "convection"? Give examples as applied to heating rooms. Enumerate the various methods of warming rooms with which you are acquainted, and describe their advantages and disadvantages.

4. Describe the advantages or disadvantages of open grates, closed stoves, and gas stoves as heating appliances for a schoolroom.

5. Describe what happens when a fire is lighted in a room with door, windows, and ventilators closed, and state what follows when a ventilator near the ceiling in the wall farthest from the fire is opened in that room.

6. What are draughts? Give diagrams showing how draughts may arise in a class-room under various conditions, and particularly with all openings, doors, and windows closed.

7. The air in a brewery and a workshop contains similar large percentages of carbon dioxide gas. Will they be equally detrimental to health? State your reasons fully.

8. Briefly describe the structure of the lungs, and explain how the purification of the blood in the lungs is effected.

9. In what way may milk convey infectious disease? and what precautions can be taken against infection in the house?

10. What in your opinion are the best methods of disinfecting (a) bedding, (b) a class-room, (c) books and boots?

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